

DOCUMENT RESUME

ED 052 964

SE 011 384

TITLE	SMIP Chemistry Curriculum Guides.
INSTITUTION	Luzerne County Schools, Wilkes-Barre, Pa.; Wilkes Coll., Wilkes-Barre, Pa.
SPONS AGENCY	Bureau of Elementary and Secondary Education (DHEW/OE), Washington, D.C.
PUB DATE	[71]
GRANT	OEG-3-7-703554-4964
NOTE	172p.
EDRS PRICE	EDRS Price MF-\$0.65 HC-\$6.58
DESCRIPTORS	*Chemistry, Curriculum Guides, *Instructional Materials, Resource Materials, *Science Activities, *Secondary School Science, *Teaching Guides
IDENTIFIERS	ESEA Title III, Science Math Improvement Project

ABSTRACT

Included are most guides for a one-year course in senior high school chemistry. The guides may be interchanged at the teacher's discretion, following any text sequence or course outline. Each guide consists of six sections: (1) an approach, which briefly discusses the unit in terms of background material, pitfalls to be avoided, and suggested methodology; (2) educational objectives, which outline the student behavior expected from successful teaching of the unit; (3) a content section, which outlines the course content to be covered; (4) a learning experience section, with suggested learning experiences such as demonstrations, lab experiments, and specialized modes of treatment for a particular concept; (5) a text reference section, and (6) an A-V section, with suggested motion pictures, filmstrips, etc., for the particular concept as far as is practicable. This work was prepared under an ESEA Title III contract. (PR)

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SMIP
CHEMISTRY
CURRICULUM GUIDES

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SE 011 384

The work presented or reported herein was performed pursuant to a Grant from the U. S. Office of Education, Department of Health, Education, and Welfare.

Grant No. OEG-3-7-703554-4964

Luzerne County School Board, Grantee
Wilkes Barre, Pa.

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SUPPLIER LISTS

Films
Filmstrips
Film Loops
Transparencies
Lab Equipment

INTRODUCTION

USING THE SMIP CHEMISTRY CURRICULUM GUIDES

The SMIP staff has tried to produce a curriculum guide that may be used by every chemistry teacher: not necessarily the entire work, possibly only several parts. In any curriculum guide, however, decisions must be made concerning content, emphasis, format, and approach. The SMIP staff feels that since no curriculum guide could satisfy every teacher unless it were written by that individual, these guides are constructed so units or subunits may be used if applicable. Even the national alphabet programs such as CBA and Chem Study are often modified to some degree by the teacher in order to fulfill some need.

For the greatest flexibility, the year's work suggested here has been divided into many small units. For the most part, the units may be interchanged in almost any fashion at the teacher's discretion, following any text sequence or course outline. Furthermore, it is not anticipated that all the material presented here could be covered in a year's introductory course. Some schools have chemistry only five periods a week, other six or seven. Some have double lab periods while others have single lab periods. A number of SMIP units are quite comprehensive, so that the entire unit could probably be covered only with a maximum time allotment, such as the SMIP unit on organic chemistry.

The length of the content outline does not necessarily indicate the amount of time to be spent in class. For example, on p. CC700-1, the early history of chemistry is covered from the Egyptians up to the modern chemists. All this material could be studied in one-half a class period up to two class periods, depending upon the depth of coverage given by the instructor. In other cases, certain parts of the content outline might be too difficult for "average" students; this material might be omitted. The time utilization of the SMIP curriculum guides is dependent upon the teacher.

Each curriculum guide consists of six sections: (1) an approach, which briefly discusses the unit in terms of background material, pitfalls to be avoided, and suggested methodology; (2) educational objectives, which outline the student behavior expected from successful teaching of the unit; (3) a content section, which outlines the course content to be covered; (4) a learning experience section, with suggested learning experiences such as demonstrations, lab experiments, and specialized modes of treatment for a particular concept; (5) a text reference section, with references keyed to the three most popular texts utilized in the Region M area; (6) an A-V section, with suggested motion pictures, filmstrips, etc. for the particular concept as far as is practicable. To the knowledge of the SMIP staff, these are the most comprehensive curriculum guides in chemistry ever assembled.

Two sections of the curriculum guides may be used to improve instruction in chemistry in this manner: duplicate the objectives and course content outline, and pass out to the students before each unit is covered in class. Thus the student knows exactly what is expected of him for each unit.

Interspersed throughout the curriculum guide are references to other SMIP publications written to supplement the units. These publications may be secured by writing to the SMIP office. A list will be found appended at the end of the curriculum guides.

In many respects, a curriculum guide is a viable thing. It must be changing as curriculum needs and trends also change. The SMIP staff welcomes constructive criticism and feedback on the guides.

GENERAL OBJECTIVES

COGNITIVE DOMAIN

- I All matter is composed of units called fundamental particles; under certain conditions these particles can be transformed into energy and vice versa.
- II Matter exists in the form of units which can be classified into hierarchies of organizational levels.
- III The behavior of matter in the universe can be described on a statistical basis.
- IV Units of matter interact. The bases of all ordinary interactions are electromagnetic, gravitational, and nuclear forces.
- V All interacting units of matter tend toward equilibrium states in which the energy content (enthalpy) is a minimum and the energy distribution (entropy) is most random. In the process of attaining equilibrium, energy transformations or matter transformations of matter-energy transformations occur. Nevertheless, the sum of energy and matter in the universe remains constant.
- VI One of the forms of energy is the motion of units of matter. Such motion is responsible for heat and temperature and for the states of matter: solid, liquid, and gaseous.
- VII All matter exists in time and space and, since interactions occur among its units, matter is subject in some degree to changes with time. Such changes may occur at various rates and in various patterns.

AFFECTIVE DOMAIN

- I Science proceeds on the assumption, based on centuries of experience, that the universe is not capricious.
- II Scientific knowledge is based on observations of samples of matter that are accessible to public investigation in contrast to purely private inspection.
- III Science proceeds in a piecemeal manner, even though it also aims at achieving a systematic and comprehensive understanding of various sectors or aspects of nature.
- IV Science is not, and will probably never be, a finished enterprise, and there remains very much more to be discovered about how things in the universe behave and how they are interrelated.
- V Measurement is an important feature of most branches of modern science because the formulation as well as the establishment of laws are facilitated through the development of quantitative distinctions.

PSYCHOMOTOR DOMAIN

- I Science apparatus is merely a logical evolution of simple tools necessary to perform certain tasks.
- II Scientists, fledgling or experienced, should be able to construct or assemble apparatus suitable for the experimental process necessary, and to use it with sufficient skill in order to secure the desired information.

SUPPLIER LIST

(Including films, filmstrips, film loops, transparencies, lab equipment)

La Pine Scientific Company
South Buckhout Street
Irvington-on-Hudson
New York, New York 10533

Cambosco Scientific Company, Inc.
342 Western Avenue
Boxton, Massachusetts 02135

Central Scientific Company
237 Sheffield Street
Mountainside, New Jersey 07092

Wilkesn-Anderson Co. (semimicro)
4525 West Division Street
Chicago, Illinois 60651

Coronet Films
Coronet Building
Chicago, Illinois 60601

Fisher Scientific Company
Gulph Road
King of Prussia, Pa. 19406

Welch Scientific Company
609 West 51st Street
New York, New York 10019

McGraw-Hill Films
330 West 42nd Street
New York, New York 10036

Audio-Visual Division (FOM)
Popular Science Publishing Co., Inc.
355 Lexington Avenue
New York, New York 10017

Association Films, Inc. (Yale Films)
Ridgefield, New Jersey 07657

Modern Learning Aids
Division of Modern Talking Picture Service, Inc.
1212 Avenue of the Americas
New York, New York 10036

Northeastern Pennsylvania Audio-Visual
Library and Instructional Materials Center
Mr. John Arcangelo, Director
Court House Annex
Scranton, Pennsylvania 18503

E. H. Sargent and Co.
35 Stern Avenue
Springfield, New Jersey 07081

INTRODUCTION TO CHEMISTRY

THE APPROACH

Many instructors utilize a short historical overview of chemistry to develop some understanding of the evolution of chemistry from simple questioning through philosophic argument to the empiricism of today. The student can be brought from the first man who wondered about fire to today when man is concerned about fundamental particles of the atom; from men who experimented for years in order to discover the mechanics of a simple reaction that we teach in a seventh-grade general science course, to those who create new compounds with complexities that stagger the imagination.

Students should realize that the study of chemistry is a human endeavor, and as such, is subjected to human frailties throughout its long history. Students should know that chemistry has been affected by superstition, stupidity, stubbornness, pride, greed, intellectual dishonesty, and many other human frailties. In spite of all this travail, the search for truth has managed to emerge constantly.

The early contributions of the Egyptian artisans and Greek philosophers are a part of chemistry's heritage that should be exposed to every student. Every student should be exposed to the original derivation of the word chemistry, although many authorities themselves may differ in its origin. (It has been stated to be of Egyptian or Greek origin, meaning either "the black art" or "the Egyptian art.")

The methods of the scientist should be covered in some detail. Brandewein's discussion in his sourcebook, "Teaching High School Science: A Book of Methods" is excellent. Call it process - scientific method - problem-solving, or whatever, the way of the scientist should be intertwined with teaching every day of the chemistry course. Chemical logic should not be covered in one class period and forgotten the rest of the year.

Students from the Region "M" area can be reminded that a replica of Joseph Priestley's laboratory may be found at Priestly House in Northumberland Borough, along Route 11 near Danville. Visitors, especially those who are presently engaged in an introductory course in chemistry, are usually astounded to see the type of apparatus used by the early greats in the field.

OBJECTIVES

1. The student should be able to state an acceptable definition of chemistry.
2. The student should be able to point out basic weaknesses of the early Greek approach.
3. The student should be able to recall the role played by the alchemists in the over-all development of the science of chemistry.

4. The student should be able to recall why Antoine Lavoisier is called "The Father of Modern Chemistry" and why his contributions to the development of chemistry as a science had such a profound influence on modern chemistry as it has developed today.
5. The student should be able to apply the scientific process to the solution of a given problem-situation, defining each particular step in the process.
6. The student should be able to state three examples of how inductive reasoning is used to solve a problem.

INTRODUCTION TO CHEMISTRY

CONTENT	LEARNING EXPERIENCES
A. Early History of Chemistry 1. Contributions of Egyptians 2. Contributions of Greeks - atomic theory 3. Contributions of Romans - more practical uses 4. The Alchemists a. Searched for Philosophers Stone b. Searched for Elixir of Life c. Kept practical arts alive 5. Natural Philosophers - beginnings of modern chemists 6. Modern Chemists a. Antoine Lavoisier b. Joseph Priestley c. Humphrey Davy	The students would be interested in knowing that the word "chemistry" comes from an old Egyptian word which means "the black art." The contributions of the Greeks and Alchemists could be used as the basis for themes or reports. Fisher Scientific Co. provides, at reasonable cost, a series of prints of ancient chemists that are quite good and make an interesting lesson in themselves. These prints may also be permanently displayed in either classroom or laboratory. Stress that Lavoisier is considered "The Father of Modern Chemistry" because he made it quantitative and relied on experiment.
TEXT REFERENCES	AUDIO-VISUAL AIDS
<u>Crucibles: The Story of Chemistry</u> by Bernard Jaffe <u>The Nature of the Universe</u> by Lucretius, Penguin Books, 1951 <u>The Origins of Scientific Thought</u> by G. de Santillana, Mentor, 1961 <u>A Short History of Chemistry</u> by Isaac Asimov, Doubleday-Anchor, 1965 <u>The Evolution of Science</u> by Metraux and Crouzet, Mentor, 1963 <u>Understanding Chemistry</u> by L. P. Lessing, Signet, 1959 <u>A Brief History of Science</u> by A. R. Hall and M. B. Hall, Signet, 1964 Baker, Bradbury, et al. pp. 9-10 Metcalf, Williams, Castka, pp. 1-5 Smoot, Price, Barrett, pp. 1-6	<u>FILMSTRIPS</u> What Is Chemistry? McGraw-Hill

INTRODUCTION TO CHEMISTRY

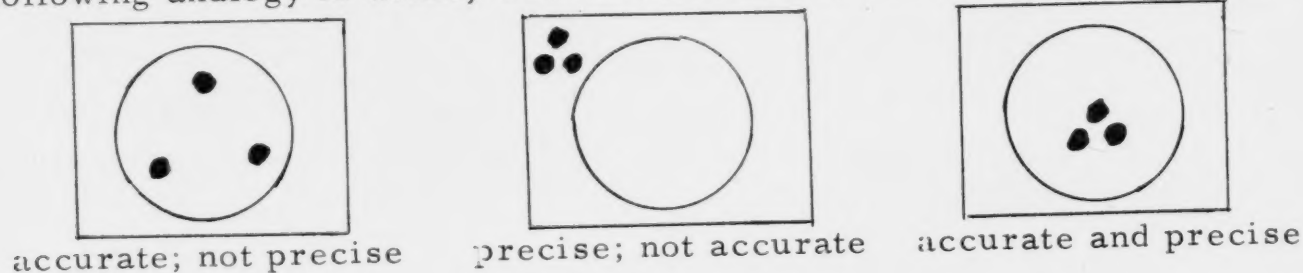
CONTENT	LEARNING EXPERIENCES
B. The Way of the Scientist 1. Classified body of knowledge a. Technical vocabulary b. Conventions in methods of expression 2. Systems of Operation a. Inductive reasoning - particular to general b. "Scientific Method" 1. Problem-solving 2. Empiricism - controls 3. Key operations 4. Concept - seeking 5. Intellectual honesty	Many teachers prefer to discuss this topic by considering "scientific method" as a set procedure consisting of these orderly steps: (1) observation (2) hypothesis or theory (3) experiment (4) conclusion. Others take the view that any logical method is valid if it works. We suggest that both approaches be discussed, along with the discussion espoused by Brandwein et al. The concept of empiricism and control should be stressed. Hypothetical examples of situations should be given to the students, who can then be asked to devise methods of operation for solving the problem
TEXT REFERENCES	AUDIO-VISUAL AIDS
1. The section entitled "Ways of the Scientist" should be considered mandatory reading. See "Teaching High School Science: A Book of Methods," Brandwein et al., Harcourt, Brace and World, 1958, pp. 11-35 2. <u>Breakthroughs in Chemistry</u> by Peter Wolff, Signet, 1967 3. <u>Chemical Engineering</u>, Nuffield #CBB04334, Penguin Books	

MATHEMATICAL OPERATIONS IN CHEMISTRY

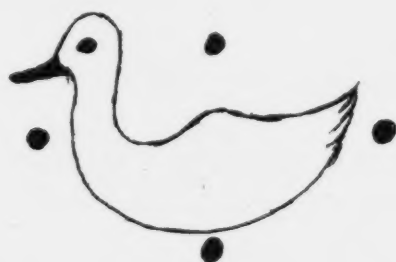
THE APPROACH

This unit--although optional in some schools--is considered by others to be one of the cornerstones of an elementary science course. When the concepts of accuracy and precision, unit analysis, and significant figures are coupled with the slide rule, the mathematical tone of the course can be set for the remainder of the year.

The opening remarks could be a short discourse showing how man's progress has been directly related to his ability to make more accurate and more precise measurements. Accuracy and precision may be interwoven into this discourse. The following analogy is widely used to demonstrate accuracy and precision:



This short analogy is worth a period of sophisticated lecture on the two concepts. Another analogy which stimulates the thinking of the student is this one:



On the average, the duck is dead.
These two simple analogies can be used to set the stage for a more vigorous treatment of experimental measurements, including Gaussian distributions and significant figures.

SMIP Handout #CC214, "Significant Figures," may be reproduced and passed out to the class, or used as the basis for a lecture/discussion on significant figures. This lesson is meaningless unless the teacher constantly enforces the rules of significant figures in all future mathematical operations! Significant figures are extremely important in physical science.

The use of unit analysis (dimensional analysis) as an aid in solving problems and securing correct answers cannot be overemphasized. However, if the teacher does not set a good example in using procedures of unit analysis and significant figures all year long in all teacher-performed computations, then the students will become lazy and forget these procedures also.

The use of the slide rule should be demonstrated, preferably with an overhead projector or a large demonstration slide rule. Constant teacher-use of the slide rule will encourage similar use among students. When students are shown that a person using a slide rule can save more than 75% of his time in solving problems

most will buy their own rules. The teacher should be somewhat lenient in marking papers and problems during the first few weeks while the students are acquiring proficiency in use of the slide rule. However, such leniency should not extend to significant figures.

Any scientific measurement involves three basic concepts: a number, a unit (dimension), and a label. An example would be if one measures the volume of nitrogen gas as 3 liters. The measurement involves the three basic concepts: a number (3), a unit (liters), and a label (nitrogen). One concept would be valueless unless the other two were present. The number 3 does not have any significance when applied to the measurement of this particular volume of nitrogen. 3 what? 3ml? 3 l? When the unit has been established, it still has no significance unless the label nitrogen is attached. This is extremely important when solving problems by the factor-label method.

The units and label of a number are part of the total quantity or measurement and must be carried along with the number. The units and label are subject to the same mathematical operations as the numbers. The units and label are complete algebraic identities.

OBJECTIVES

1. The student should be able to differentiate between the term accuracy and the term precision.
2. The student should be able to indicate how many significant digits appear in a given number.
3. The student should be able to apply the rules for the proper number of significant figures in each of the following types of operations involving scientific measurements:
 - a. multiplication
 - b. division
 - c. addition
 - d. subtraction
4. The student should be able to solve seven out of ten conversion problems using the factor-label method.
5. The student should be able to solve nine out of ten problems from each of the following types using a slide rule:
 - a. multiplication
 - b. division
 - c. square root
 - d. cube root (optional)

MATHEMATICAL OPERATIONS IN CHEMISTRY

CONTENT	LEARNING EXPERIENCES
A. Significant Figures 1. Accuracy of measurements 2. Precision of measurements 3. Rules for significant figures B. Unit Analysis (Dimensional Analysis) 1. Need for number, unit, and label 2. Carry all units and labels 3. Perform same operations on units and labels as numbers C. Use of Slide Rule 1. Assistance given by slide rule a. Aids in understanding significant figures b. Saves time 2. Multiplication (C & D scales) 3. Division (C & D scales) 4. Square Root (A & D scales) 5. Cube Root (K & D scales)	A-1 Discuss difference between accuracy and precision. Show how uncertainty in measurement makes it necessary to use concept of significant figures. B. Discuss importance of carrying units and labels along with numbers. Discuss exceptions - pure numbers like SpGr. C. Demonstrate use of slide rule, preferably linear type. SMIP Handout #CC205, "Shortcut to the Slide Rule," may be reproduced and passed out to students.
TEXT REFERENCES	AUDIO-VISUAL AIDS
Baker, Bradbury et al., pp. 54-58 Metcalfe, Williams, Castka, pp. 18-19 Smoot, Price, Barrett, pp. 23-26, 28	Many organizations produce slide rules for overhead projection, both the linear and circular type. Among these are the 3M Company, with an accompanying unit containing a projection slide rule and a set of projection masters on its use and the utilization of significant figures.

THE METRIC SYSTEM

THE APPROACH

It is generally assumed that students are able to make measurements of simple properties of matter - such as mass and length - from their earlier experiences in elementary school and junior high; but actually few know how to make accurate and precise measurements of any properties of matter. Facility in making, recording, and interpreting measurements - along with the interpretation of data from such processes - should be secured from a thorough study of this unit and the previous unit. Proper principles of measurement should be stressed in this section and in all the succeeding sections given in the course. However, the teacher should not act fanatically about proper measurement in the early months of the course and then forget about it entirely in the latter parts of the year.

The lesson can be introduced with a short discussion of some early units of measurement for matter, such as the cubit, the grain, etc. followed by the necessity for standards of measurement. Students should realize that any unit may be used as a standard. For example, a book may be picked up and said to be the new standard for mass, and the students are instructed to compare the masses of other objects in the room against the standard mass of that particular book. A short discussion of the advantages and disadvantages of the use of that particular book as a standard could lay the groundwork for understanding other standards, such as the "standard kilogram" and the distance between the two scratches on the old "standard meter."

The difference between fundamental units and derived units could be discussed in conjunction with three fundamental quantities of mass, length, and time. Many students consider volume, area, etc. to be fundamental units.

The difference between the metric system and the British system of measurement can be shown vividly by comparing the monetary systems of the U. S. A. and England. Even Australia has recently changed their monetary system (similar to England's) to a decimal system like that of the U. S. A. Learning the metric units can be facilitated by comparing the subdivisions of the dollar with the metric prefixes.

1	dollar = standard unit
$\frac{1}{10}$	dollar = <u>deci</u> dollar = dime
$\frac{1}{100}$	dollar = <u>centi</u> dollar = cent
$\frac{1}{1000}$	dollar = <u>milli</u> dollar = mill (used in taxation)

Some teachers make it mandatory for the students to memorize all the metric prefixes and use them in conversions. This is unnecessary; only the more common ones are ever used. Students who are facile in the use of unit and label analysis should have no difficulty in making conversions from one factor of ten to the other.

Conversion factors may be taught more easily through discovery in the laboratory. If the students are given objects to measure using both the metric and British systems, they should be able to construct their own conversion factors, or at least realize the order of magnitude differences between the various metric and British units. It is possible to secure weights, rules, and graduated cylinders in the British system for use in such an experiment.

OBJECTIVES

1. The student should be able to define each of the following units of measurement:
 - a. mass
 - b. weight
 - c. length
 - d. time
 - e. volume
 - f. area
 - g. work
2. The student should be able to state the basic units of length, mass, time, and volume in each of the following systems of measurement:
 - a. metric (mKs)
 - b. metric (cgs)
 - c. English
3. The student should be able to define each of the following prefixes used in the metric system:
 - a. kilo-
 - b. centi-
 - c. milli-
 - d. micro-
4. The student should be able to solve twenty-five out of thirty problems dealing with metric conversions which involve mass, length, and volume.

THE METRIC SYSTEM

CONTENT	LEARNING EXPERIENCES
A. Measurement and Man's Progress 1. Early units of measurement 2. Evolution of measuring tools and increase in precision 3. Standards a. Primary b. Secondary B. Fundamental Units of Measurement 1. Mass (and weight) 2. Length 3. Time C. Derived units of measurement 1. Volume 2. Area 3. Work 4. All other units	A. Assign class collateral reading from science periodicals, such as <u>Science World</u>, old editions of <u>SMW</u>, <u>Science</u>, etc. concerning the metric system. Show how some systems of measurement are easier to use than others. Show how any system of measurement can be made up from a primary standard. Select a book, call it a unit of mass, and compare other masses against it. B-C The student should be shown how all other units can be derived from the three fundamental units, such as $\text{volume} = \text{length} \times \text{length} \times \text{length}$. If the instructor deems it necessary, the distinction between mass and weight can be made. Compare weights and masses on the earth and moon.
TEXT REFERENCES	AUDIO-VISUAL AIDS
Baker, Bradbury, et al., pp. 50-51 Metcalfe, Williams, Castka, pp. 11-13 Smoot, Price, Barrett, pp. 14-16	<u>EBF Filmstrip</u> <u>MEASURING DISTANCE</u> 11.122-C-Si Excellent introduction in formation of measuring systems. Very elementary. First half should be shown with this lesson; other half on use of metric distances. <u>SVE Filmstrip</u> <u>EXPERIMENTS WITH LENGTH</u> 449-1 <u>EXPERIMENTS WITH MASS</u> 449-2

THE METRIC SYSTEM

CONTENT	LEARNING EXPERIENCES
D. The Metric System 1. Decimal system 2. Easily compared standards E. Metric Prefixes 1. Mega- (M) = 1,000,000 x (not common) 2. Kilo- (k) = 1,000 x 3. Hecto- (h) = 100 x (not common) 4. Deka- (dk) = 10x(not common) 5. Deci- (d) = 0.1 (not common) 6. Centi- (c) = 0.01 7. Mili- (m) = 0.001 8. Micro- (μ) = 0.000,001	Open lesson with discussion of USA monetary system (100¢ = \$1.00; 10¢ = 1 dime, etc.) Show how both U.S. monetary system and metric system are both based on the same concept. Students should memorize the more common prefixes. Drill on conversions by moving the decimal point.
TEXT REFERENCES	AUDIO-VISUAL AIDS
Baker, Bradbury, et al., pp. 50-51 Metcalfe, Williams, Castka, pp. 11-13 Smoot, Price, Barrett, pp. 32-33	<u>EBF Filmstrip</u> 1121C-Si THE METRIC SYSTEM OF MEASURING Excellent elementary treatment of the metric system and conversions into the British system. <u>FOM Filmstrip</u> SCIENTIFIC MEASUREMENT - MOLECULES TO STARS (588) Range of scientific measurements and magnitudes explored. Precision is stressed. Using very simple equipment, students may reproduce the work of Dr. Langmuir.

THE METRIC SYSTEM

CONTENT	LEARNING EXPERIENCES
F. Metric Units 1. Length - the meter 2. Mass - the gram 3. Time - the second 4. Volume - the liter G. Conversion Factors 1. Length 2. Mass 3. Volume	Drill on metric units, their multiples, and conversions into the British system. Show students how dimensional analysis and significant figures may be used to make conversions easier. Students should be encouraged to use the metric system exclusively in the chem room - similar to language teachers demanding use of foreign languages only in advanced language studies. At least thirty different problems involving metric conversions of mass, length, volume, including problems deriving volume from length and area, could be assigned for homework.
TEXT REFERENCES	AUDIO-VISUAL AIDS
Baker, Bradbury, et al., pp. 50-51 Metcalfe, Williams, Castka, pp. 11-13 Smoot, Price, Barrett, p. 33	<u>EBF Filmstrip</u> MEASURING WEIGHT Elementary treatment of the difference between mass and weight, and the difference between balances and scales.

DENSITY AND SPECIFIC GRAVITY

THE APPROACH

This unit is a logical outgrowth of the previous unit on the metric system. The concept of varying densities of substances can be quite effectively shown with a few simple demonstrations, especially the one in which four immiscible liquids of different densities are kept in the same hydrometer jar.

Dimensional analysis should be stressed throughout this unit; if dimensional analysis is accepted and utilized, the density formula need not be used. Density and SpGr again demonstrate the role of a standard in chemistry.

Specific gravity should be treated simply as a ratio of two densities; introduction of Archimedes' Principle and treatment of SpGr in the manner of the physics course need not be introduced at this point. It should be stressed that SpGr is a pure number and thus does not have units.

OBJECTIVES

1. The student should be able to define the term density.
2. The student should be able to solve nine out of ten of each of the following types of problems:
 - a. determine density of a solid or liquid, given the mass and volume
 - b. determine mass, given the density and volume
 - c. determine volume, given the density and mass
3. The student should be able to define the term specific gravity.
4. The student should be able to solve nine out of ten specific gravity problems dealing with solids, liquids, and gases

DENSITY AND SPECIFIC GRAVITY

CONTENT	LEARNING EXPERIENCES
A. Necessity for Volumetric Measurement of Liquids 1. Inconvenient to weigh 2. Weight determinations time-consuming B. Definition - mass per unit volume $D = \frac{M}{V}$ C. Density Units 1. Solids and liquids a. metric - g/ml b. British - lb/ft³ 2. Gases - g/liter 3. Density standard - water at 4° C D. Density Problems	A. Demonstrate or discuss difficulty of weighing liquids. B. Exhibit one pound bottles of mercury, silver nitrate, calcium acetate, aluminum oxide. Show that each bottle contains one pound, yet each bottle has a different volume. C. Secure identical cubic specific heat specimens from physics department or construct identical cubic specimens from different materials. Weigh each, calculate volume from linear measurements, and compute density of each.
TEXT REFERENCES	AUDIO-VISUAL AIDS
Baker, Bradbury, et al., p. 159 Metcalf, Williams, Castka, pp. 6-8 Smoot, Price, Barrett, pp. 17-18	<u>FILM LOOPS</u> VOLUME (EBF) No. 20603-Color; 3 min. By means of animation and live photography, the concept of volume as a three-dimensional measure is developed. Illustrates basic volume units used in chemical laboratory practice. DENSITY (EBF) No. 20602-Color; 3 min. Presents an actual derivation of the density of a solid together with the qualitative measure of liquid density. Defines density as weight/unit volume. Illustrates the relationship between the density of a solid and a liquid. <u>FILMSTRIPS</u> MEASURING VOLUME AND DENSITY(EBF) No. 11120 Although prepared for intermediate grades, this is an excellent treatment for chemistry.

DENSITY AND SPECIFIC GRAVITY

CONTENT	LEARNING EXPERIENCES
A. Definition - ratio of density of unknown to density of standard 1. Solids and liquids $\text{SpGr} = \frac{D \text{ unknown}}{D \text{ water}}$ 2. Gases $\text{SpGr(A)} = \frac{D \text{ unknown}}{D \text{ air}}$ B. Advantages of SpGr - no units, pure number; can be changed into any density system. C. Specific Gravity Problems	Prepare a cylinder containing four immiscible liquids with different specific gravities; add Hg, CCl_4, H_2O, and petroleum ether in that order to a hydrometer jar. If dye is needed, add methyl red powder to each liquid. <u>Do not shake up layers</u> as petroleum ether and CCl_4 will mix, resulting in three layers. Keep for next year. Show that wood, cork and iron nut added to above device will sink to different levels because of SpGr of each.
TEXT REFERENCES	AUDIO-VISUAL AIDS
Baker, Bradbury, et al., pp. — Metcalf, Williams, Castka, pp. — Smoot, Price, Barrett, pp. —	Many excellent A-V materials prepared primarily for physics can be shown concerning specific gravity. However, in chemistry the simplified treatment of SpGr given above does not need a discussion of SpGr from the standpoint of Archimedes' Principle. Discussing Archimedes' Principle at this time only complicates the point at hand and consumes time which may be spent more fruitfully on other concepts. <u>FILMSTRIPS</u> (Jam Handy Density and Specific Gravity Physics Series) SPECIFIC GRAVITY OF SOLIDS AND LIQUIDS

COMPOSITION OF MATTER

THE APPROACH

Much of the material covered in this unit is descriptive and may have been covered in some earlier science courses; it may be presented as a rapid review to reinforce the concepts and possibly correct any misconceptions.

Broadly speaking, this unit should not only show that matter is discontinuous and consists of finite, discrete, ultimate particles, but also should introduce the student to some of the important vocabulary and definitions which will be used frequently later in the program. Facility in a technical vocabulary is highly desirable, and the subject matter unit can lay the groundwork for such a vocabulary in chemistry.

Demonstrations can be used to great advantage in showing the difference between elements, compounds, mixtures, and chemical/physical changes. Time can be saved for later use if the teacher utilizes demonstrations rather than lab experiments; however, good open-ended experiments in classifying matter may be a better choice.

OBJECTIVES

1. The student should be able to define the term matter.
2. The student should be able to indicate why matter possesses the property of inertia.
3. The student should be able to explain the definite relationship which exists between matter and energy.
4. The student should be able to state and interpret the Law of Conservation of Matter-Energy.
5. The student should be able to indicate the three states in which matter can exist.
6. The student should be able to explain the differences which exist among the solid, liquid, and gaseous states in terms of molecular activity and distance between molecules.
7. The student should be able to define the term substance.
8. The student should be able to give ten examples of substances.
9. The student should be able to distinguish between a homogeneous substance and a heterogeneous substance.
10. The student should be able to define the term element.
11. The student should be able to explain the fundamental difference between a metal, a non-metal, and a metalloid.
12. The student should be able to give ten examples of an element.
13. The student should be able to define the term compound.
14. The student should be able to solve nine out of ten problems dealing with the percentage composition of compounds in terms of the Law of Definite Composition.

15. The student should be able to define the Law of Definite Proportions and correctly interpret the law using a given chemical equation such as $\text{Fe} + \text{S} \rightarrow \text{FeS}$.
16. The student should be able to give ten examples of a compound.
17. The student should be able to define the term mixture.
18. The student should be able to point out the basic difference between a mixture and a compound.
19. The student should be able to give ten examples of a mixture.
20. The student should be able to differentiate between an atom and a molecule.
21. The student should be able to differentiate between a physical property and a chemical property.
22. The student should be able to list ten physical properties of a substance.
23. The student should be able to list five chemical properties of a substance.
24. The student should be able to identify physical changes and chemical changes.
25. From a list of chemical and physical properties, the student should be able to identify correctly nine out of ten as being either a physical property or a chemical property.

COMPOSITION OF MATTER

CONTENT	LEARNING EXPERIENCES
A. Matter 1. Definition - anything that occupies space 2. Property of inertia 3. Energy and matter B. States of Matter 1. Solid - definite shape and definite volume 2. Liquid - definite volume and indefinite shape 3. Gas - indefinite volume and indefinite shape	A-1 Display different common elements in identical glass vials, such as S, C, Sn, Sb, I₂, Mg, Al, Ca, Zn, Br₂. A-2 Discuss inertia by means of analogies. A-3 Define energy, work, force. B. Discuss states of matter from definitions given at left; use block of wood, two different vessels of odd shape to hold liquids, rubber balloon for gases.
TEXT REFERENCES	AUDIO-VISUAL AIDS
Baker, Bradbury, et al., pp. 10-11 Metcalfe, Williams, Castka, p. 28 Smoot, Price, Barrett, pp. — <u>Chemicals and Where They Come From</u> Nuffield #CBB 04313, Penguin Books	<u>EBF Filmstrip</u> CLASSIFICATION OF MATTER

COMPOSITION OF MATTER

CONTENT	LEARNING EXPERIENCES
C. Discontinuity of Matter 1. Substance - a particular kind of matter a. Homogeneous b. Heterogeneous 2. Element - the simplest form of matter retaining properties of the whole a. Metals b. Non-metals c. Metalloids 3. Compound - chemical union of two or more elements a. Law of Definite Composition b. Law of Multiple Proportions 4. Mixture - physical union of two or more substances	C-1 Demonstrate homogeneity with sand, mineral fragments, and pieces of wood. Show heterogeneity with muddy water. C-2 Discuss discontinuity of matter; refer to discussions of ancient Greeks. C-3 Alyea demonstration 1-5. Make Fe + S mixture; separate Fe with magnet, dissolve S with CS₂. Heat to form FeS; show formation of compound by failure of same tests. Use stoichiometric quantities. Assign for homework list of ten common elements, ten common compounds, ten common mixtures.
TEXT REFERENCES	AUDIO-VISUAL AIDS
Baker, Bradbury, et al., pp. 20-23 Metcalfe, Williams, Castka, p. 28 Smoot, Price, Barrett, pp. 41-42	<u>SVE Filmstrip</u> WHAT THINGS ARE MADE OF (427-20) ELEMENTS, COMPOUNDS, AND MIXTURES(493-2) <u>PSSC Film</u> ELEMENTS, COMPOUNDS, AND MIXTURES (Color) #0111 34 minutes A discussion of the difference between elements, compounds, and mixtures, showing how a mixture can be separated by physical means. Demonstrates how a compound can be made and then be taken apart by chemical methods with identification of components by means of their physical properties such as melting point, boiling point, solubility, color, etc.

COMPOSITION OF MATTER

CONTENT	LEARNING EXPERIENCES
D. Ultimate Particles 1. Atom - smallest particle of element which can have properties of that element. 2. Molecule - chemical union of two or more atoms-either like or unlike. If unlike atoms, then molecule is smallest particle which has properties of that compound. E. Properties of Matter 1. Physical properties - change occurs when physical form changes. 2. Chemical properties - change occurs when atoms are re-arranged and may result in composition change.	D. Recall discontinuity of matter; the name "atom" is what scientists gave to the fundamental particle of an element. Discuss atoms and elements as building blocks from which more complicated structures, molecules, are made. E. Enumerate physical properties such as density, boiling point, melting point, etc. and chemical properties such as color, reactivity, etc.
TEXT REFERENCES	AUDIO-VISUAL AIDS
Baker, Bradbury, et al., pp. 11-15 Metcalf, Williams, Castka, p. 35 Smoot, Price, Barrett, p. 45	<u>EBF Filmstrips</u> THE ATOM MAN DISCOVERS THE ATOM (#8860-9) <u>PSSC Film</u> DEFINITE AND MULTIPLE PROPORTIONS #0110 30 minutes Here is the evidence on which Dalton based his conviction that matter came in natural units, atoms; the chemical laws of definite proportions demonstrated by electrolysis and recombination of water; and multiple proportions by the quantitative decomposition of N_2O, NO, and NO_2

COMPOSITION OF MATTER

CONTENT	LEARNING EXPERIENCES
F. Changes in Matter 1. Physical changes a. Change in state b. Energy involved 2. Chemical changes a. Changes which result in new substances with new properties b. Energy involved	F-1 Discuss physical changes and examples. Demonstrate Alyea 1-6. Tear paper, break stick, melt and resolidify naphthalene. Assign ten examples of physical changes for homework. F-2 Alyea Demonstrations 1- 8 Burn Mg ribbon - flash of light 1- 9 Rusty iron bar - slow oxidation Fe fillings in burner - rapid oxidation 1-10 Ammonium dichromate volcano 1-12 Add Fe nail to CuSO_4-aq. Copper plates out. 1-14 Cut tree from Cu foil; add to AgNO_3-aq. Silver tree forms.
TEXT REFERENCES	AUDIO-VISUAL AIDS
Baker, Bradbury, et al., pp. 16-18 Metcalfe, Williams, Castka, pp. 36-37 Smoot, Price, Barrett, pp. 46-47 <u>Chemicals from Nature</u>, Nuffield #CBB 04328, Penguin Books <u>Burning</u>, Nuffield #CBB 04324, Penguin Books	<u>EBF Filmstrip</u> CHEMICAL CHANGE (8040-6) <u>SVE Filmstrip</u> CHEMICAL CHANGES (427-21)

ATOMIC STRUCTURE

THE APPROACH

The unit on atomic structure may be initiated with a discussion of the historical development of the topic. Dependent upon the class involved, it will be necessary to make a decision where the historical development of atomic structure should begin. A good starting point for many regular chemistry classes is the Thomson Model. The fundamental weaknesses of Thomson's Plum Pudding Model (Grapes-in-Jello Model) should be thoroughly discussed. It can be shown that the alpha scattering experiment of Rutherford contradicted the Thomson Model. The film entitled "The Structure of Atoms," McGraw-Hill #612022, and the PSCC film #0416 "Rutherford Atom" are useful aids and emphasize this particular phase of the development of the atom's structure. Before continuing the discussion, however, the basic weaknesses of the Rutherford atom should be pointed out.

Sequential development may be logically continued with a discussion of the discovery of the electron and proton using the cathode ray tube experiment as the basis of discussion. As a result of this discussion, the Bohr theory of the atom can be introduced. It should be shown that the Bohr Atom overcame certain existing weaknesses which were prevalent relative to the structure of atoms up to this time. At this point, depending upon the ability and background of the class, it may be helpful to show how Bohr used the work of Planck to explain quantization of energy levels in the atom. An interpretation of the equation $E = h\nu$ and how it led to Bohr's conclusions should be covered.

Before continuing beyond this point, it should be pointed out that the Bohr Atom has certain basic weaknesses and these weaknesses should be discussed. Following consideration of the discovery of the neutron (1932), the modern view of nuclear structure may be introduced. A brief discussion of nuclear structure may be continued dealing with such topics as mass defect and binding energy.

A discussion of the electronic structure of the atom may now be undertaken. The FOM filmstrip #616 "Atomic Structure and Chemistry" and the Yale film loop YF #251 "Electronic Structure of Elements" are valuable aids to use in proving the existence of basic atomic particles and to introduce the correlation between atomic structure and basic quantum theory. Probability distributions and the charge cloud concept should be covered. An attempt should be made to correlate quantum numbers with shells, subshells, orbitals, and electron spin. A valuable aid at this point is the FOM filmstrip #629 "Orbitals—Atomic and Molecular."

A thorough discussion of the rules for assigning quantum numbers, the order of increasing energy levels, the Pauli Exclusion Principle, and Hund's Rule

should be undertaken. The paperback "Atoms, Crystals, and Molecules" Part 1 by Ainslie Drummond is excellent for clarifying the above-mentioned topics. A helpful hint to memorize the increasing order of energy levels is as follows:

Write the following groups of numbers and note the pattern:

12
23
34
345
456
4567
5678

These numbers represent the energy shells in the atom (principle quantum number). Each time a number appears the first time, assign an "s" subshell; each time a number appears the second time, assign a "p" subshell; each time a number appears a third time, assign a "d" subshell, etc. For example, if the above pattern of numbers is rewritten and we follow the above rules, the following is obtained:

1s2s
2p3s
3p4s
3d4p5s
4d5p6s
4f5d6p7s
5f6d7p8s

It would now be appropriate and timely to discuss stable and unstable arrangements of electrons in orbitals within subshells. The film "The Structure of Atoms" #1165 NE Pa. AV Center is effective for this purpose. Electron configurations may be related to the structure of the periodic table, indicating that metals, nonmetals, metalloids, transition elements, and the rare-earths and actinides are positioned in the periodic table according to their electron configurations. The periodic table should be covered in greater detail (trends in properties, etc.) in the unit dealing with the periodic table.

Before completing the discussion of atomic structure, isotopes should be discussed. Effective tools for this topic are the film "Isotopes" #1091 NE Pa. AV Center, the SMIP #1 filmstrip "Properties of Isotopes," and the McGraw-Hill film "Isotopes."

OBJECTIVES

1. The student should be able to state the contributions made by the early Greeks which eventually led to the modern theory of atomic structure.
2. The student should be able to state the basic principles of Dalton's atomic theory.
3. The student should be able to give a brief historical account of the discovery of the fundamental atomic particles—electron, proton, and neutron.
4. The student should be able to explain the significance of Rutherford's Gold-Foil Experiment which elucidated the structure of the atomic nucleus.
5. The student should be able to explain the contributions of Bohr's theory to atomic spectra and atomic structure.
6. The student should be able to state the following atomic properties based upon the Bohr model:
 - a. relative size
 - b. relative density
 - c. charge
 - d. electron motion about the nucleus
 - e. axial rotation
7. The student should be able to determine the number of neutrons, protons, and electrons in an atom, given the atomic number and the mass number.
8. The student should be able to state the similarities in chemical properties and the differences in physical properties of isotopes of a given element.
9. The student should be able to give four quantum numbers of any electron in a given atom.
10. The student should be able to show an understanding of the Pauli Exclusion principle by assigning quantum numbers to electrons.
11. The student should be able to determine the maximum number of electrons in each energy shell and subshell of an atom.
12. The student should be able to show the correlation between quantum numbers and the following structural concepts of the atom:
 - a. energy shell
 - b. energy subshell
 - c. orbital
 - d. electron spin
13. The student should be able to state and illustrate Hund's Rule of maximum multiplicity.
14. The student should be able to state the order of increasing energy levels in an atom.

ATOMIC STRUCTURE

CONTENT	LEARNING EXPERIENCES
A. <u>Introduction</u> 1. Greek concept of indivisible particles a. Atoms b. Discontinuity of matter 2. Dalton's atomic theory a. Derivation of atomic theory b. Postulates of atomic theory	A. Develop historical background with student reports on Democritus, Dalton, and/or Berzelius. Discuss Dalton's atomic theory as an explanation for the Law of Definite Composition and the Law of Multiple Proportions.
TEXT REFERENCES	AUDIO-VISUAL AIDS
Baker, Bradbury, <u>et al.</u>, pp. 24-25 Metcalf, Williams, Castka, pp. 41-46 Smoot, Price, Barrett, pp. 104-110 <u>Atomic Structure & Bonding</u> Basic Systems/Lyons & Carnahan, p. 7. Good programmed text. <u>Microstructure of Matter</u> USAEC Understanding the Atom Series.	NE PA A-V CENTER AUDIO TAPES SC-270 All about atoms—Dalton and the Atomic Theory - 15 min. SC-175 Dalton Finds the Atom - 14-1/2 min. LIFE FILMSTRIPS - The Atom (55 frames) McGRAW-HILL FILMSTRIP - Atomic Theory

ATOMIC STRUCTURE

CONTENT	LEARNING EXPERIENCES
B. <u>The Nuclear Atom</u> 1. Fundamental particles a. Protons c. Neutrons b. Electrons 2. Rutherford experiment a. Nucleus 1) Small 3) Positively charged 2) Dense 4) Motion-spin on axis b. Extra-nuclear electrons 1) Distance from nucleus much larger than diameter of nucleus 2) Density-only $\frac{1}{1840}$ proton 3) Charge-negative to positive protons 4) Motion a) Spins on axis b) Travels around nucleus c) Speed 0.9 that of light	B1. Discuss discoveries of fundamental particles and compare and contrast their masses and charges. Discuss significance of symbols ${}_{-1}e^0$, ${}_1p^1$ (${}_1H^1$), and ${}_0n^1$. B2. Discuss significance of Rutherford Gold-Foil Experiment in leading to concept of nuclear atom. Discuss how Bohr modified the nuclear theory of Rutherford with his planetary explanation of electron configuration. Stress that this is a "far-away look" at the atom; a closer look will show finer structures, e.g. such as energy levels, etc. Stress also that these theories are only man-made explanations to explain natural phenomena.
TEXT REFERENCES	AUDIO-VISUAL AIDS
Baker, Bradbury, <u>et al.</u>, pp. 30-35 Metcalf, Williams, Castka, pp. 46-48 Smoot, Price, Barrett, pp. 112-113 <u>Inside the Atom</u> Nuffield #CBB04330 Penguin Books Baker, Bradbury, <u>et al.</u>, pp. 32 Metcalf, Williams, Castka, pp. ——— Smoot, Price, Barrett, pp. 123-127	NE PA A-V CENTER 1334 How Big is An Atom? BW 30 min. FOM FILMSTRIP 616. ATOMIC STRUCTURE AND CHEMISTRY Properties of fundamental atomic particles are experimentally demonstrated. Periodic Table of Elements is validated and the experimental theory of atomic structure explained. Chemical properties and atomic structure are correlated and elementary quantum theory is applied. THE STRUCTURE OF ATOMS This film provides the experimental evidence for our basic concepts concerning the structure of the atom. An experiment similar to Rutherford's historic alpha particle scattering demonstration shows that atoms have dense, positively charged nuclei. Another fundamental experiment shows the charge on the electron and the ratio of charge to mass. Color (612022) 12-1/2 min. B & W(612010)

ATOMIC STRUCTURE

CONTENT	LEARNING EXPERIENCES
C. <u>Structure</u> 1. Nucleus a. Atomic number b. Atomic mass c. Mass number d. Isotopes e. Nomenclature of nuclides $\begin{matrix} Z \\ A \end{matrix} X_N$ Z = atomic number A = mass number A = Z + N N = neutron number	C. Relate concepts of atomic number, mass number, and isotopes by illustrating and comparing hydrogen isotopes, helium, and lithium. Stress dependence of change in properties of elements with atomic number and not mass number. Isotopes have same chemical properties because of same Z, but slightly different physical properties. Show SMIP Filmstrip #1, "Properties of Isotopes."
TEXT REFERENCES	AUDIO-VISUAL AIDS
Baker, Bradbury, <u>et al.</u>, pp. 34-35 Metcalfe, Williams, Castka, pp. 47 Smoot, Price, Barrett, pp. 115-120	McGRAW-HILL FILMSTRIP Set #2 019005 Structure of the Atom NE PA A-V CENTER 1091 Isotopes C 15 min. 1165 The Structure of Atoms C 12 min. FOM FILMSTRIP 645. EVOLUTION OF THE ELEMENTS Discusses the structure of the atomic nucleus from the historical viewpoint and in terms of nucleons, concepts of atomic number, mass number, and isotopes, and two fundamental processes involved in nucleogenesis and nucleosynthesis. Discusses cosmic abundance of elements; uses two nuclear processes to explain the production and relative abundance of the elements heavier than iron. SMIP FILMSTRIP #1 "Properties of Isotopes" ISOTOPES - The film shows uranium being separated into two isotopes-U-238 & U-235. It explains how J. J. Thompson first demonstrated the existence of isotopes and how Aston developed the first mass spectrometer. It then shows two methods of separating isotopes and concludes by illustrating the uses of radioisotopes. C 15 min.

ATOMIC STRUCTURE

CONTENT	LEARNING EXPERIENCES
C. (continued) 2. Electronic energy levels a. The Bohr atom b. Orbital concept 1) Principal energy levels: 1, 2, 3, ... 2) Sublevels: s, p, d, ... 3) Orbitals c. Maximum number of electrons in energy levels = $2n^2$ 1) Sublevels = number of principal level 2) Orbitals, s = 1, p = 3, d = 5, f = 7	C2. Discuss electromagnetic radiation including Planck's concept and atomic spectra as background for the concept of the Bohr atom. Note the necessity of energy sublevels to explain the multitude of observed atomic spectral lines.
TEXT REFERENCES	AUDIO-VISUAL AIDS
Baker, Bradbury, <u>et al.</u>, pp. 195-206 Metcalf, Williams, Castka, pp. 51-60 Smoot, Price, Barrett, pp. 144-153	A transparency may be made from the Fisher-Cook Atomic Orbital Board (Fisher Scientific Company #5-702-10) and the electron buildup demonstrated with an overhead projector. The students can practice writing electron configurations using the Fisher-Cook Orbital Board.

ATOMIC STRUCTURE

CONTENT	LEARNING EXPERIENCES
D. <u>Electron Configurations</u> 1. Order of increasing energy 2. Hund's rule of maximum multiplicity 3. Valence electrons 4. Atomic kernels E. <u>Quantum numbers</u> 1. Principal: $n = 1, 2, 3 \dots$ 2. Subsidiary: $\ell = 0, \dots, n-1$ 3. Magnetic: $m = +1 \dots 0 \dots -1$ 4. Spin: $s = \pm 1/2$	D. Note increasing energy of electrons with increasing distance of probability and density from nucleus. Include concept of Hund's rule as additional energy needed to place like charges together. Illustrate order in which electrons enter various orbitals stressing maximum number of electrons that each level and sublevel can contain and the order of increasing energy. Note valence electrons for each schematic developed. E. Discuss Pauli Exclusion principle and its relationship to individual sets of quantum numbers for electrons. Relate quantum numbers to maximum numbers of electrons that can be contained in each subshell and shell. Diagrammatically show that the third energy level can contain only 18 electrons by illustrating the possible combinations for $n = 3$
TEXT REFERENCES	AUDIO-VISUAL AIDS
Baker, Bradbury, <u>et al.</u>, pp. 195-206 Metcalfe, Williams, Castka, pp. 51-60 Smoot, Price, Barrett, pp. 144-153 Baker, Bradbury, <u>et al.</u>, pp. 207-219 Metcalfe, Williams, Castka, pp. 54-55 Smoot, Price, Barrett, pp. 144-153	YF-251 Electronic Structure of Elements C—11 min. FOM FILMSTRIP 629. ORBITALS—ATOMIC AND MOLECULAR Sets of four quantum numbers serve to provide the basis for various schematics showing the electronic configuration of elements. Overlap of atomic orbitals results in formation of molecular orbitals of the sigma and pi types. Resultant molecular geometries are illustrated.

CHEMICAL BONDING

THE APPROACH

It will be necessary early in the discussion of chemical bonding to emphasize the role which valence plays in bond formation. The film "Chemical Bond and Atomic Structure" by Coronet Films may be helpful in accomplishing this initial objective.

Ionization potential, electron affinity, and electronegativity should be defined such that the part they play in predicting the type of bond formed can later be understood more fully. (FOM Filmstrip #648 "Ionization Potential and Electronegativity.")

It should be mentioned that there are basically three types of chemical bonds: ionic or electrovalent, covalent, and metallic. It should then be demonstrated and shown that tables of ionization potentials, electron affinities, and electronegativities can be used to predict the type of bond that will be formed between two given atoms. (FOM Filmstrip #627 "Covalent Bonds - Covalent Structures.")

Each of the three basic types of chemical bonds can now be discussed separately as individual lessons. The FOM Filmstrip #6505, "The Ionic Bond" and Filmstrip #6506, "The Covalent Bond: Molecular Orbitals" are effective aids which may be used at this point. The coordinate covalent bond can now be introduced as a special type of covalent bond. The FOM Filmstrip #6507, "The Coordinate Covalent Bond" is an excellent teaching device to introduce and explain this topic. In order to develop an understanding of a metallic bond the FOM Filmstrip #666, "Metals and the Metallic Bond" is quite effective.

It should be pointed out that bond formation is the result of the minimum potential energy (maximum stability) requirement of atomic systems. Also it may be shown how the energy of the system involving two atoms diminishes as the atoms move closer to each other, until a minimum potential energy is attained. If the atoms are brought even closer together, the potential energy will increase and repulsion will occur. A simple plot of potential energy versus distance will vividly illustrate this phenomena.

After each of the three basic bond types have been discussed, an attempt should be made to crystallize the overall picture of bond formation. The fundamental prerequisites for the formation of each bond type should be discussed. Molecular model sets are effective aids and should be used especially when discussing covalent bonding. Before concluding the covalent bond presentation, it will be necessary to introduce and explain the single, double, and triple covalent bond. A comparison of bond strength and of each type should also be included in the presentation.

The concept of resonance and hybrid bonds should be covered before concluding the unit on bonding. Evidence for the existence of both should be given.

OBJECTIVES

1. The student should be able to state the octet rule and interpret the stability of noble gases with respect to the rule.
2. The student should be able to show how the following atomic properties help to determine the combining capacity of an atom:
 - a. electronegativity
 - b. electron affinity
 - c. ionization potential
 - d. electron configuration
3. The student should be able to give three examples of electron transfer reactions.
4. The student should be able to write the oxidation half-reaction, the reduction half-reaction, and the overall net-ionic equation to illustrate the formation of a given ionic bond.
5. The student should be able to state the meaning of the term empirical formula using the sodium chloride crystal as an example.
6. The student should be able to write Lewis dot structures for five ionic compounds.
7. The student should be able to give three examples of reactions leading to covalent bonds.
8. The student should be able to write a Lewis dot structure for a molecule containing a single covalent bond, a molecule containing a covalent double bond, and a molecule containing a covalent triple bond.
9. The student should be able to write a Lewis dot structure for a coordinate covalent compound such in NH_4^+ , and state the difference between a coordinate covalent bond and an ordinary covalent bond.
10. Given a table of electronegativity values, the student should be able to predict the type of bond expected from a combination of atoms in a given compound.
11. The student should be able to write the formulas of five molecules which have polar covalent bonds.
12. The student should be able to write the formulas of five molecules which have a non-polar covalent bond.
13. The student should be able to write the formulas of five molecules or radicals which are said to possess hybrid bonds.

CHEMICAL BONDING

CONTENT	LEARNING EXPERIENCES
A. Formation of a Chemical Bond 1. Relationship of valence electron configuration to combining capacity of atoms a. attainment of stable noble gas type octets b. prediction of combining ratios 2. Transfer of valence electrons (ionic bond) a. formation of ions b. determination of ionic charges 3. Sharing of valence electrons (covalent bond)	A-1 Stress the tendency of atoms to attain inert gas type configurations; derive from electronic configurations the formulas for a series such as NaCl, Mg, Cl₂, AlCl₃. A-2 Demonstrate that both electron transfer and electron sharing reactions may occur by discussing the formation of NaCl and H₂ from atoms.
TEXT REFERENCES	AUDIO-VISUAL AIDS
<u>Chemical Bonds: Introduction and Fundamentals</u>, Eichinger, J. W., Lyons & Carnahan (Good programmed learning text.) <u>Atomic Structure & Bonding, Basic Systems</u>/Appleton-Century-Crofts, Lyons & Carnahan, p. 243 (Very good programmed materials on bonding)	<u>Coronet Films</u> CHEMICAL BOND AND ATOMIC STRUCTURE (16 min.) This film discusses the structure of the atom and shows that the electrons in the outermost shell, or energy level, of an atom help determine the way in which it bonds chemically with other atoms. Animation and laboratory demonstrations show three types of chemical bond and explain how bonding affects the physical and chemical properties of a substance. <u>McGraw-Hill Filmstrip</u> CHEMICAL BOND, Set No. 3, #019100 <u>FOM Filmstrip</u> HYDROGEN - THE FUNDAMENTAL ATOM, #622 Occurrence of hydrogen and its importance to man are shown in an examination of its uses in our technology. Unique properties of hydrogen are discussed, correlated with its atomic structure. Use of hydrogen in nuclear research and technology is explained.

CHEMICAL BONDING

CONTENT	LEARNING EXPERIENCES
B. Bonding Extremes 1. Ionic (electrovalent) bonds a. transfer of valence electrons i. formation of cations by loss of electrons ii. formation of anions by gain of electrons b. compound formation by attractions of oppositely charged ions i. arrangement of ions in a compound ii. significance of empirical formulas	B-1 Illustrate, using models, the change that occurs during formation of an ionic bond such as NaCl in terms of: 1. $M^0 \longrightarrow M^{+1} + e^{-}$ 2. $Y^0 + e^{-} \longrightarrow Y^{-1}$ 3. $M^{+1} + Y^{-1} \longrightarrow M^{+}Y^{-}$ Display model of NaCl crystal to demonstrate ionic arrangements and to stress significance of an empirical formula.
TEXT REFERENCES	AUDIO-VISUAL AIDS
Atoms, Crystals, Molecules, Parts I and II, Drummond, Ainslie H., American Education Publications (Excellent supplementary text.)	<u>FOM Filmstrip</u> THE IONIC BOND, #6505 Deals with the disperse nature of ionic bonds, the empirical formula, unit cell and lattice structure, and related radius ratios. Characteristic properties and variation in bond strengths particularly as they affect melting points and hardness are considered in detail. THE HALOGENS (Second Edition) (16 min.) The highly reactive halogens combine with metallic elements to form ionic compounds, and with non-metallic elements to form covalent compounds. Uses such as silver bromide on film and newer developments such as Teflon coating are shown. The combining of fluorine with noble gases is also illustrated. Animation and laboratory demonstrations show how atomic structure is related to the high reactivity of halogen elements and to the stability of their compounds. (Color-\$180; Black and white-\$90; Senior High or College)

CHEMICAL BONDING

CONTENT	LEARNING EXPERIENCES
B-1 (Continued) c. energy changes involved in electron transfer i. ionization energy of metal ii. electron affinity of non-metal iii. ordering of ions in a compound iv. resulting net energy release (exothermic) d. Lewis Dot Symbolism for ionic compounds e. effect of ionization on atomic radii i. decrease in radius with cation formation ii. increase in radius with anion formation	Relate energy changes to steps in formation of an ionic compound. 1. $M^0 + \text{Ionization Energy} \rightarrow M^{+1} + e^-$ 2. $Y^0 + e^- \rightarrow Y^{-1} + \text{Electron Affinity Energy}$ 3. $M^{+1} + Y^{-1} \rightarrow M^+Y^- + \text{Attraction Energy}$ 4. $M^0 + Y^0 \rightarrow M^+Y^- \quad E = \frac{\text{Electron}}{\text{Affinity}} + \frac{\text{Attraction}}{\text{Energy}}$ d. Show students how to write Lewis dot structures for common atoms and ions. e. Display chart on atomic and ionic radii (Welch [CHEM] #4844) and discuss generalizations on effect of ionization of radii.
TEXT REFERENCES	AUDIO-VISUAL AIDS
Baker, Bradbury, et al., pp. 222-223 Metcalf, Williams, Castka, p. 81 Smoot, Price, Barrett, pp. 194-202	

CHEMICAL BONDING

CONTENT	LEARNING EXPERIENCES
B. Bonding Extremes (continued) 2. Covalent bonds - sharing of valence electrons to complete octet - overlap of atomic orbitals to form bond - relationship of molecular formula to atomic electron disturbances - comparison of molecular and ionic units	Illustrate covalent single, double, and triple bonds by displaying molecular models of H_2, O_2, and N_2. Use common ball-and-stick models to illustrate single, double, and triple bonds in compounds like C_2H_6, C_2H_4, C_2H_2.
TEXT REFERENCES	AUDIO-VISUAL AIDS
Baker, Bradbury, et al., pp. 227-229 Metcalf, Williams, Castka, pp. 85-86 Smoot, Price, Barrett, pp. 188-194	<u>FOM Filmstrips</u> THE COVALENT BOND: MOLECULAR ORBITALS #6506 Presents the formation of molecular orbitals from atomic orbitals and their use in dealing with the elements from hydrogen through neon. Molecular orbitals as well as hybridization of orbitals are used to explain the structure of more complex compound molecules. COVALENT BONDS-COVALENT STRUCTURES #627 Briefly reviews the structure of electrovalent compounds. Types of covalent bonds and resulting covalent structures are related to the concepts of electronegativity and electronegativity differences. Variety of models are used in the development of concepts. MACROMOLECULAR BONDS #6508 Explains the fundamental characteristics of such network solids as diamond, silicon dioxide, and graphite and others in terms of the strong covalent bonds present in the three dimensional aggregate. Properties indicating the strength of this type of bonding include hardness, high melting point, and conductivity.

CHEMICAL BONDING

CONTENT	LEARNING EXPERIENCES
B-2 (Continued) c. energy change involved in electron sharing i. liberation of energy upon attainment of noble gas electron configuration ii. energy requirement for separating diatomic molecules d. coordinate covalent bonds e. Lewis dot symbolism for molecular compounds f. radicals i. covalency ii. similarity to single ions 3. Metallic Bonds a. nature of "electron gas" b. bonding by positive ion-electron cloud attractions	d. Illustrate coordinate covalency by discussing the bonds in the ammonium ion and contrasting these to the N-H bonds and electron distribution in the ammonia molecule.
TEXT REFERENCES	AUDIO-VISUAL AIDS
Baker, Bradbury, et al., pp. 231-234 Metcalf, Williams, Castka, pp. 85-88 Smoot, Price, Barrett, pp. 194-204 <u>The Structure of Molecules</u>, Barrow, Gordon M., W. A. Benjamin, (Teacher Reference) <u>Metals, Atoms, and Alloys</u>, McCabe, C. L., and Bauer, C. L., NSTA, Vistas of Science Series	<u>FOM Filmstrips</u> THE COORDINATE COVALENT BOND, #6507 Investigates the role of coordinate covalent bond formation of polyatomic ions, certain molecules, and Werner complexes. Utilizes related concepts such as the valence bond theory, orbital hybridization, and the ligand field theory. METALS AND THE METALLIC BOND, #666 Outlines the general physical properties of metals, with experiments indicating the relative mobility of the valence electrons. Develops the theory of the metallic bond and illustrates representative crystalline structures of metals. Explores such metallic properties as characteristic ionization energies and oxidation potentials. METALS OF GROUPS IA AND IIA, #659 These highly electropositive elements have typical metallic properties. Their high chemical reactivity is related to metallic bonding, ionization energies, and electronegativities.

CHEMICAL BONDING

CONTENT	LEARNING EXPERIENCES
C. Bond Polarity 1. Electronegativity a. Definition b. Pauling electronegativity scale i. low values of metals ii. high values of non-metals iii. values of metalloids iv. relationship between electronegativity values and activity c. Relationship of electronegativity difference to bond type	Discuss unit with the aid of the typical electronegativity chart found in most text books or other charts. Show how an electropositivity scale can be formed with the same set of units. Discuss the relation of electronegativities to naming of compounds composed of non-metals such as S_2Cl_2. Show how electronegativity difference can be used as guide to predict bond type; use KI, CO, $SbCl_3$ as examples.
TEXT REFERENCES	AUDIO-VISUAL AIDS
Baker, Bradbury, et al., pp 220-222 Metcalfe, Williams, Castka, pp. 91-93 Smoot, Price, Barrett, pp. 198-199	<u>Yale Film Loop</u> (Chem Series) ELECTRONEGATIVITY, #YF-244 (4 min. - Color) <u>FOM Filmstrip</u> IONIZATION POTENTIAL AND ELECTRONEGATIVITY, #648

CHEMICAL BONDING

CONTENT	LEARNING EXPERIENCES
C. Bond Polarity (continued) 2. Polar covalent bonds - effect of comparative electronegativities - directional character - linearity - non-linearity D. Hybrid Bonds - Evidence for existence - unexpected combining capacities - unexpected bond equalities - Reasons for formation - promotion and hybridization energy requirement - stability from additional bond energies	Discuss lack of polarity of bonds in H_2, O_2, and N_2 in terms of a zero electronegativity difference. Develop concept of polar bonds by discussing electronegativity differences of elements in HCl, $BrCl$, H_2O, and NH_3. Discuss problem of relating the formulas and bond equalities of simple carbon compounds, such as CH_4, CCl_4, and CO_2 to the electron configuration of the carbon atom.
TEXT REFERENCES	AUDIO-VISUAL AIDS
Baker, Bradbury, et al., pp. 229-230 Metcalf, Williams, Castka, p. 92 Smoot, Price, Barrett, pp. 199-200 <u>Coordination Chemistry</u>, Basolo, F. and Johnson, R., W. A. Benjamin (Teacher Reference)	<u>Chem Study Film</u> SHAPES AND POLARITIES OF MOLECULES #4154 (18 min. - Color) Electrical effects including a stream of falling liquid by an electrically charged rod develop the concept of molecular polarity. Shows a model based on polar and non-polar molecules. Molecular dipole model is extended to explain differences in solubility, conductivity, and chemical reactivity.

THE PERIODIC TABLE

THE APPROACH

An introduction to the study of the periodic table through historical background can make the evolution of the periodic table like a scientific detective story. SMIP has devised a discovery experiment or learning experience to simulate the struggles of the early workers in formulating the periodic table. (SMIP #CC213) If this experiment is used to introduce the unit on the periodic table, it should give the students insight into the early struggles of Mendeleeff, Meyer, Newlands, et al., and lay the groundwork for the historical treatment.

When discussing Mendeleeff's contribution to the periodic table, the assumption of error in atomic weights should be discussed thoroughly as a classic example of "forcing the piece with the right patterns to fit."

A simple open-ended learning experience has been constructed for students which allows appreciation of some of the problems of the early chemists in formulating a periodic table; called "The Periodic Puzzle", it consists of a number of blocks of fictional elements and chemical and physical data for each. The student is given the puzzle to cut out and rearrange in any way he chooses as long as his construction shows the periodicity of the elements and trends in their properties. He must also predict the properties of an unidentified element from an empty space in the puzzle. The exercise may give the student a greater understanding of the role played by the periodic table in chemistry.

As the unit progresses, studies should be undertaken on trends of chemical properties, such as atomic radius, ionization energy, metallic and non-metallic character, electronegativity, etc. At the end of the unit, the student may be asked to perform certain tasks from the list of learning objectives given for this unit. The ability to fill in a blank periodic table is essential; the student can demonstrate his knowledge of the location of various families and groups by doing this.

OBJECTIVES

1. The student should be able to recall three of the early attempts to classify elements.
2. The student should be able to state the weaknesses of the early classifications.
3. Given Mendeleeff's table, the student should be able to give the atomic number, mass number, approximate density, state of matter, and color (if solid), of a given element in terms of its position in the table.
4. The student should be able to interpret chemical and physical property trends of elements from one member of a group to the next member

- and from one member of a series to the next member.
5. The student should be able to identify elements as either metals, non-metals, or metalloids from their position in the periodic table.
 6. The student should be able to predict the location of element number 105 in the table and to predict at least five properties of that element.
 7. The student should be able to determine a possible substitute for an element such as aluminum in products containing aluminum.
 8. The student should construct some type of periodic table from SMIP #CC213, "The Periodic Puzzle."

ACTIVITIES BASED ON THE PERIODIC TABLE

Many generalizations summarized below may be verified from data available in well-known reference books. This type of activity serves to give the pupil experience in the use of reference materials, a skill required in advanced sciences. Two excellent sources are Handbook of Chemistry and Physics, published by the Chemical Rubber Publishing Company, and Handbook of Chemistry by N. A. Lange, published by McGraw-Hill.

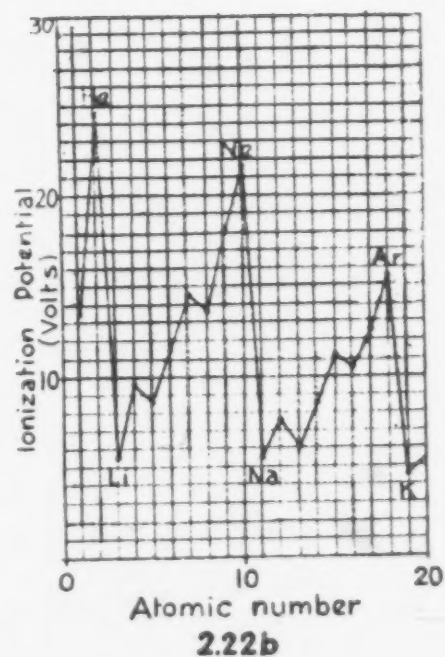
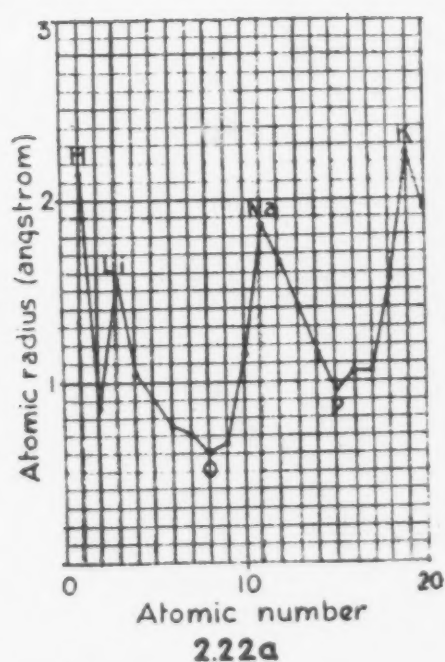
If desired, data may be obtained for elements other than those suggested in this and in the preceding activity. The explanations of various interesting exceptions to the generalizations may be beyond the scope of an introductory course.

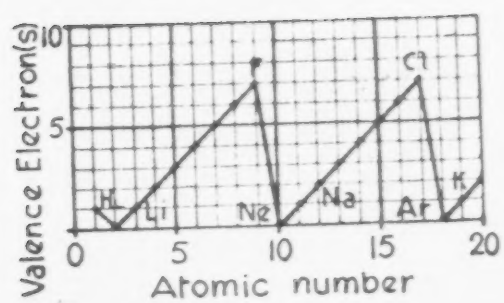
Many of the obvious activities require constant reference to the periodic table. Trace on a master duplicating ditto sheet pertinent portions (squares, atomic numbers, period and group numbers) of the Periodic Table of the Elements distributed. Reproduce in sufficient quantities to allow each pupil enough copies to record on separate sheets various data (some of these data may also be plotted as illustrated in diagrams 2.22a, b, c, and d), such as:

- (1) The boiling points of elements by group, period or both (note particularly Groups IA, IIA, VIIA and O).
- (2) The melting point of elements by group, period or both.
- (3) The density of elements by group, period or both.
- (4) The atomic radii of the elements (refer to Lange under "Atoms, effective radii of"). Note that data from various sources differ slightly.
- (5) The ionic radii of the elements (refer to Campbell's Chart of Relative Sizes of Atoms and Ions in the Periodic Table or to Lange as above). Note that data from various sources differ slightly.
- (6) The first ionization potentials of the elements by group, period or both.

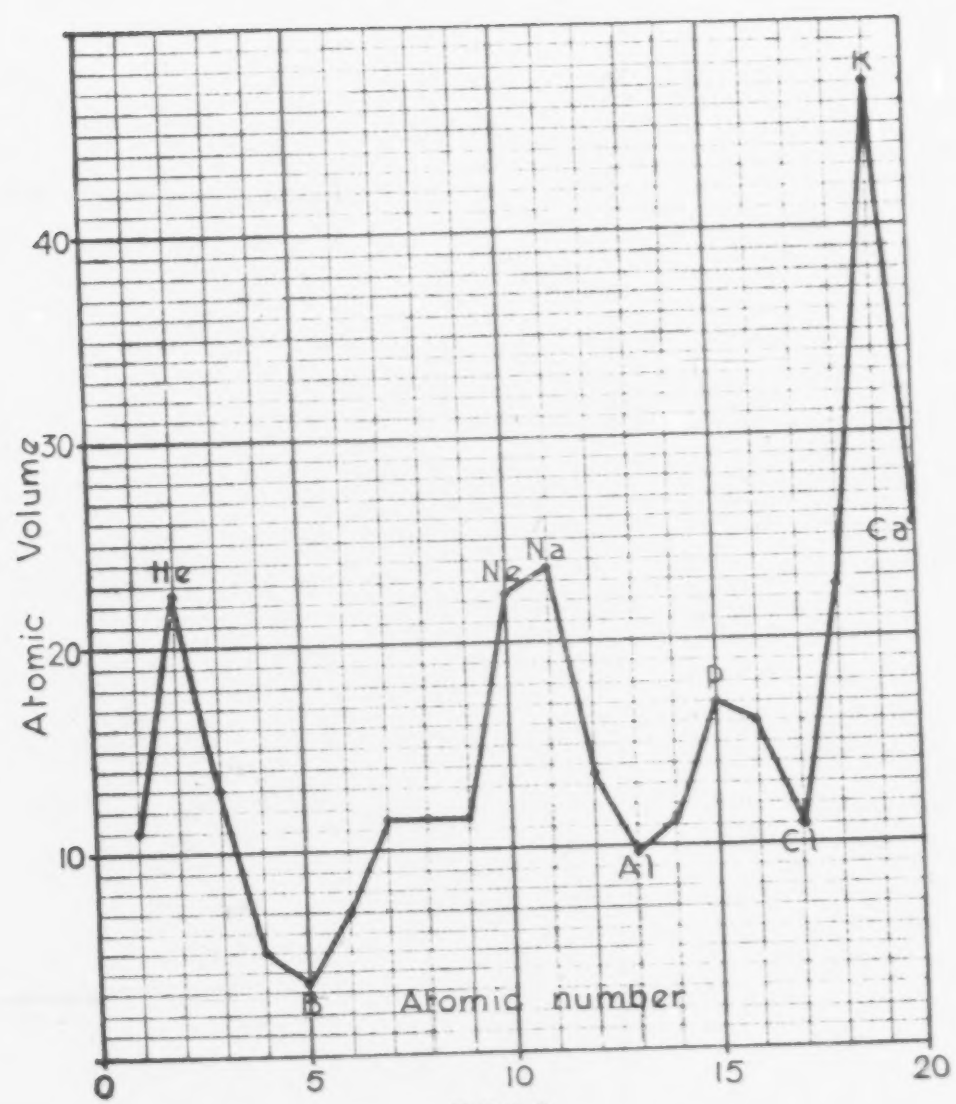
Adapted from "Chemistry Handbook," published by Bureau of Secondary Curriculum Development, N. Y. State Education Department, Albany, 1962.

- (7) The electronegativities of the elements(exclude the transition elements).
- (8) The melting points of a representative sample of ionic compounds.
- (9) The boiling points of a representative sample of covalent compounds.
- (10) The number of valence electrons for the elements with atomic numbers 1 through 20.
- (11) The atomic volume of the number of elements with atomic numbers 1 through 20.





2.22c



2.22d

PERIODIC TABLE

CONTENT	LEARNING EXPERIENCES
A. Necessity for classification 1. Difficulty of remembering all characteristics of all substances. 2. Allows one to predict properties of an element from position in periodic table. 3. Demonstrates order in chemistry B. Early attempts at classification 1. Lavoisier's groups 2. Dobereiner's triads 3. Newland's octaves	A-1 Illustrating data in "Handbook of Chemistry." Discuss impossibility of remembering all information in the book. B Discuss early attempts to classify elements. If time permits, assign students "Periodic Table Puzzle" for solving in one class period.
TEXT REFERENCES	AUDIO-VISUAL AIDS
Baker, Bradbury, et al., pp. 189 Metcalfe, Williams, Castka, pp. 64-65 Smoot, Price, Barrett, pp. 158-159 "The Periodic Table," Chem BG Book, Nuffield Project, Penguin Books, 1967. (Excellent short unit in the periodic system.) <u>The Chemical Elements</u>, Nuffield # . CBB04331, Penguin Books. (Good historical treatment of some elements.) Read the life of Mendeleyeff in "Crucibles," by Jaffe (pocket edition).	<u>Chem Study Film</u>, # 4112 CHEMICAL FAMILIES Starting with a display of actual samples of over 70 elements, the film demonstrates methods by which chemical similarities among the elements have provided the basis for dividing them logically into families. The elements are classified as metals or non-metals, and some are found to be difficult to classify. Experimentally it is shown that, under the same conditions, some cases are chemically reactive, while others are not. The fact that elements with atomic numbers one less and one more than the atomic numbers of the inert gases are highly reactive, provides the clue for finding the halogens and alkali metal families. The film shows how atomic numbers have provided the basis for ordering of the elements in the periodic table. (22 min. - Color)

PERIODIC TABLE

CONTENT	LEARNING EXPERIENCE
C. Solving the Puzzle 1. Meyer's curves Periodicity 2. Mendeleyeff's table a. Showed periodicity b. Predicted undiscovered elements c. Based on Atomic weight I and Te interchange 3. Moseley's Atomic Numbers Resolved irregularities in Mendeleyeff's table D. Use of the Periodic Table 1. Prediction of properties from orderly trends 2. Led to new discoveries	C-1 Have class plot atomic volume versus atomic weight of alkali metals and halogens to show periodicity. Or show transparency of graph. C-2 Discuss periodicity: Show how gaps and Mendeleyeff's predictions clarified the puzzle. Discuss irregularities and "logical explanations" which were set forth and accepted. C-3 Show transparency of Moseley's plots of properties versus Z and compare with plots versus A from Mendeleyeff.
TEXT REFERENCES	AUDIO-VISUAL AIDS
Baker, Bradbury, et al., pp. 190-192 Metcalfe, Williams, Castka, pp. 66-72 Smoot, Price, Barrett, pp. 160-164 <u>The Start of X-ray Analysis</u>, Nuffield #CBB04327 <u>Atomic Structure and Bonding</u>, Basic Systems/Lyons & Carnahan, p. 63 (Good programmed unit.)	<u>McGraw-Hill Filmstrip</u> THE FAMILY OF HALOGENS The film shows the similarity and gradation of properties of the halogens, and explains both the similarities and differences in degree on the bases of atom structure. Valuable in developing the general use of the Periodic Table. (12 1/2 min. - Color[Code 612025] B & W [Code 612000]) <u>FOM Filmstrip</u> IONIZATION POTENTIAL AND ELECTRO-NEGATIVITY, #648 Provides an understanding of the meaning of ionization potential, electron affinity, and electronegativity. Magnitudes and periodicity related to atomic structure, macroscopic chemical properties and bonding.

PERIODIC TABLE

CONTENT	LEARNING EXPERIENCES
E. Statement of Periodic Law 1. Mendeleyeff version 2. Moseley version F. Arrangement of Modern Periodic Table 1. Elemental blocks a. Chemical symbol b. Atomic number c. Atomic weight (mass) d. Electronic configuration 2. Periods: Horizontal a. Variance from strong metallic to strong non-metallic properties with increasing number of outer shell electrons. b. Close of period with noble gas configuration (octet)	Student must be cautioned that Moseley's version is <u>more</u> tenable than the first Mendeleyeff version. F. Exhibit periodic table and discuss features of each elemental block. Use transparencies and the overhead projector for this unit. F-2 Referring to periodic table, relate electronic configurations to variance in metallic-non-metallic properties of elements moving across periods 2 and 3. Note close of each period with a noble gas.
TEXT REFERENCES	AUDIO-VISUAL AIDS
E-F-1: Baker, Bradbury, et al., pp. 192-194 Metcalf, Williams, Castka, pp. 66-72 Smoot, Price, Barrett, pp. 164-170 <u>Introductory Descriptive Chemistry</u>, by Ronald C. Johnson. W.A. Benjamin (Teacher Reference) F-2: Baker, Bradbury, et al., pp. 208-219 Metcalf, Williams, Castka, pp. 66-70 Smoot, Price, Barrett, pp. 164-169	<u>McGraw-Hill Filmstrip</u> PERIODIC TABLE, Set #2 019005 <u>Welsh Scientific Company</u> # 39925, Chemistry for Today: Group II The Periodic System <u>FOM Filmstrip</u> THE CHEMISTRY OF PERIOD 3, #625 Major emphasis on variation in properties of the elements in the sequence: highly reactive metal-metals of decreasing activity - metalloids - non-metals of increasing activity - noble gas. Hydrides and oxides are treated in detail.

PERIODIC TABLE

CONTENT	LEARNING EXPERIENCES
c. Transition series i. Filling of inner electron shells ii. Similarity of properties 3. Groups (Families): Vertical Columns a. Similarity of properties for elements of similar outer shell electron configurations b. Francium: Most active metal i. One outer electron in all Group I elements ii. Most complex member: Electron most easily removed	Locate transition series on the table. Stress filling of inner electron shells and constant number of outer electrons as the series proceed 3. Discuss similar properties of alkali metals (Group I), halogens stressing relationship of characteristic properties for each family to common valence and electron configuration. Student experiment to compare and contrast properties of Cl_2, Br_2 and I_2.
TEXT REFERENCES	AUDIO-VISUAL AIDS
Baker, Bradbury, et al., p. 217 Metcalf, Williams, Castka, pp. 66-70 Smoot, Price, Barrett, pp. 169-170; 178-182	Northeastern Pennsylvania Audio-Visual Center VANADIUM - A TRANSITION ELEMENT, #A-836 (Color - 22 min.) FOM Filmstrip THE TRANSITION ELEMENTS, #663 The approximately 30 metals comprising the transition elements in the Periodic Table are thoroughly discussed to explain paramagnetic and ferromagnetic behavior, to demonstrate multiplicity of and changes in oxidation states, and to develop understandings of ions and catalytic activity. Northeastern Pennsylvania Audio-Visual Center CHEMICAL FAMILIES, #A-832, (Color - 22 min.) THE FAMILIES OF HALOGENS, #1045, (Color - 12 min.) THE DOSIUM FAMILY (16 min.) The physical and chemical properties of the alkali metals are related to their atomic structures, and their compounds are shown through laboratory demonstrations, including the rarely performed reaction between sodium and chlorine. The Downs Cell, Solvay process, salt formation by neutralization, water softening and flame tests are included, in addition to a number of important new uses.

PERIODIC TABLE

CONTENT	LEARNING EXPERIENCES
F-3 (Continued) c. Fluorine: Most active non-metal i. Seven outer electrons in all Group VII elements ii. Simplest member: greatest attraction for one electron - highest electronegativity d. Group VIII - Noble Gases i. Electron octets ii. Tendency to inert behavior toward chemical reaction e. Subgroups i. A and B families ii. Similarity of outer shell configurations of A and B elements iii. Lack of similarity of chemical properties of A and B elements	3-c By means of charts, either student-made as homework, or teacher-prepared, show selected chemical properties which demonstrate the family concept. 3-e Compare similar valence electron configuration of Groups I-A and I-B, noting similar formulas for compounds. Contrast Groups I-A and I-B by noting difference in reactivities of the metals and different valences.
TEXT REFERENCES	AUDIO-VISUAL AIDS
Baker, Bradbury, et al., pp. 217-129 Metcalfe, Williams, Castka, p. 70 Smoot, Price, Barrett, pp. 180-181 <u>The Discovery of the Halogen Elements</u>, Klopfer, Leo G., HOSC Series/SRA. <u>Rare Earths: The Fraternal Fifteen</u> <u>Synthetic Transuranium Elements</u>, USAEC, Understanding the Atom Series	<u>FOM Filmstrip</u> THE NOBLE GASES, #647 Argon, Helium, Krypton, Neon, Radon, and Xenon--their discovery, general chemical inertness, and practical uses are presented in this colorful film-strip. The chemistry of noble-gas compounds is described in detail.

PERIODIC TABLE

CONTENT	LEARNING EXPERIENCES
G. Periodic Properties 1. Atomic radii a. Decrease across a period b. Increase down a group c. Factors affecting charge on ion i. Increase in nuclear charge ii. Increase in electron repulsions iii. Addition of an electron shell	G-1 Develop concept of atomic radii as a function of the number of occupied principal energy levels.
TEXT REFERENCES	AUDIO-VISUAL AIDS
Baker, Bradbury, et al., pp. 220-222 Metcalfe, Williams, Castka, pp. 70-72 Smoot, Price, Barrett, p. 175	<u>FOM Filmstrip</u> BORON AND ALUMINUM, #668 Presents group III-A elements in terms of electronic configurations, ionization energies, atomic radii, and electronegativities. Filmstrip covers non-metallic characteristics and nature and behavior of important compounds of boron, and contrasts the behavior of aluminum with the nonmetallic behavior of boron. The production and key properties of aluminum are also discussed.

PERIODIC TABLE

CONTENT	LEARNING EXPERIENCES
2. Ionization energies a. Definition b. Low values of metals c. High values of nonmetals d. Unusually high values of noble gases e. Decrease with increasing atomic radius within a group i. Distance of outermost electron ii. Shielding effect of inner electrons and neutralizing effect on increasing nuclear charge.	G-2 Compare the first ionization energies for the elements of Li to Ne as a function of atomic radius. Similarly compare the first ionization energies of Groups I and VII.
TEXT REFERENCES	AUDIO-VISUAL AIDS
Baker, Bradbury, et al., pp. 222-223 Metcalfe, Williams, Castka, pp. 72-75 Smoot, Price, Barrett, pp. 194-198	<u>Chem Study Films</u> IONIZATION ENERGY, #4151, (Color - 22 min.) Shows how ionization energies are experimentally determined.

PERIODIC TABLE

CONTENT	LEARNING EXPERIENCES
G-2 (Continued) f. Irregular increase across a series i. Decrease from Group II to Group III ii. Decrease from Group V to Group VI g. Ionization of successive electrons i. Increase with each additional electron removed ii. Marked change between principal electronic energy levels 4. Electron affinities a. Definition b. High values of halogens c. Relationship to ionization energy	Compare first and successive ionization energies of Na, Mg, and Al stressing relationship to orbital designation of electron(s) involved.
TEXT REFERENCES	AUDIO-VISUAL AIDS
Baker, Bradbury, et al., pp. 222-224 Metcalfe, Williams, Castka, pp. 72-76 Smoot, Price, Barrett, pp. 194-200	59

COMPOSITION CALCULATIONS

THE APPROACH

The concepts in this unit may be taught immediately after the students have learned how to balance equations, or interspersed throughout the course.

The section on formula weights can be covered early with some emphasis on the other types of nomenclature. Formula weights of hydrates should be illustrated because it will help later in the preparation of solutions. It is suggested that the Mole concept and Avogadro's number be introduced preferably early, immediately after formula weights.

Percentage composition, simple formula, and molecular formula problems should be covered at the same time because the problem types are closely related to each other. Many students do not realize that simple formula problems are really only percentage problems worked in reverse. Determination of the molecular formula of a hydrate should also be shown.

This unit contains many possibilities for quantitative experiments utilizing sensitive balances such as the Welch #4030 or Mettler if one is available.

One experiment which is excellent for illustrating quantitative techniques is the determination of the percentage of oxygen in potassium chlorate. Another similar experiment is the determination of the formula of a hydrate (such as $\text{BaCl}_2 \cdot 2\text{H}_2\text{O}$) from its percentage composition.

OBJECTIVES

1. The student should be able to determine the formula weight of any compound, given the atomic weights (mass numbers) of the elements in the compound and the simple formula.
2. The student should be able to determine the percentage composition of any given compound, given the mass number of each element in the compound.
3. The student should be able to determine the empirical formula of any given compound, given the percentage composition of that compound.
4. The student should be able to determine the molecular formula of any compound, given the percentage composition and molecular weight of the compound.
5. The student should be able to show the specific relationship which exists between Avogadro's number and the Mole concept.

6. The student should be able to solve nine out of ten problems of the types which follow:
 - a. Converting weights of compounds into moles
 - b. Converting weights of compounds into number of molecules
7. The student should be able to define an amu.
8. The student should be able to state specifically the relationship which exists between a gram and an amu.

COMPOSITION CALCULATIONS

CONTENT	LEARNING EXPERIENCES
A. Compounds 1. Formula weights 2. Percentage composition 3. Empirical formulas a. Definition b. Calculation from percentage composition	A-1 Problem assignments in calculating formula weights, percentage composition by weight, and empirical formulas. Discuss need for formula weights. A-2 Student experiment on determination of percent composition, such as the determination of percent oxygen in KClO_3. A-3 Student experiment on calculating the formula of a hydrate (such as barium chloride hydrate) by determining the percent H_2O in the compound.
TEXT REFERENCES	AUDIO-VISUAL AIDS
Baker, Bradbury, et al., pp. 44, 87, 127, 141-142 Metcalf, Williams, Castka, pp. 100-108 Smoot, Price, Barrett, pp. 75, 81-82	<u>McGraw-Hill Filmstrip</u>, Set #1 CHEMICAL FORMULA, #019000 <u>Welch Filmstrip</u>, Group I SYMBOLS, FORMULAS, EQUATIONS, ATOMIC AND MOLECULAR WEIGHTS, #3992R <u>Chemistry Lab Series</u> MOLECULAR WEIGHT OF OXYGEN The loss of weight of potassium chlorate after heating is determined and the molecular weight of oxygen is calculated. (11 min. - \$100 Color-- \$50 B/W)

COMPOSITION CALCULATIONS

CONTENT	LEARNING EXPERIENCES
A (Continued) 4. Molecular formulas a. Relationship to empirical formula b. Calculation from percentage composition and molecular weight	A-4 Student experiment on determination of the formula of a compound, such as the formula of CuS or MgO.
TEXT REFERENCES	AUDIO-VISUAL AIDS
Baker, Bradbury, et al., pp. 141-142 Metcalf, Williams, Castka, p. 109 Smoot, Price, Barrett, p. 83	NE PA A-V CENTER <u>McGraw-Hill Film</u> DETERMINING MOLECULAR FORMULAS, #1021 An experiment shows how the molecular weight and formula of a compound may be determined through knowledge of the percentage composition by weight of the compound and by the application of Avogadro's Laws. (12 1/2 min. - Color (Code 612020) \$150; B/W (Code 612008) \$75) <u>Chemistry Lab Series</u> SYNTHESIS OF A COMPOUND The simplest formula of a cuprous sulfide is derived by determining the weight ratio, then performing the conventional calculations. (13 1/2 min. - Color \$125 B/W \$62.50)

COMPOSITION CALCULATIONS

CONTENT	LEARNING EXPERIENCES
B. Mole Concept 1. Definition - Avogadro's number 2. Relative atomic weights a. Dalton's Theory b. $H = 1$ c. $O = 16$ d. $C = 12$ 3. Gram atomic weights a. Definition of amu b. Relationship of an amu to the gram 4. Gram molecular weights	B. Show how relative atomic weights are related to percentage composition of compounds. Show that relative mass data and percent weight data lead to the same empirical formula. B-3 Problem assignments on converting weight into moles and converting weights into number of molecules.
TEXT REFERENCES	AUDIO-VISUAL AIDS
Baker, Bradbury, et al. , pp. 35, 127-130 Metcalfe, Williams, Castka, pp. 104-105 Smoot, Price, Barrett, pp. 76-80	<u>FOM Filmstrip</u>, Set #A-21 THE MOLE CONCEPT, #643 The introductory frames treat the Mole concept and its relation to the Avogadro number. Related stoichiometry is extensive with emphasis on the use of the Mole concept in solving problems.

HEAT

THE APPROACH

The unit may be opened with a simple demonstration showing that a 100 ml beaker of boiling water and a 400 ml beaker of the same liquid do not cool at the same rate - their heat contents are different, although the two liquids are both at the same initial temperature. The comparison between heat and temperature should be made. Temperature may be roughly discussed as one indication of the heat content of a substance, while heat is a form of energy.

Early attempts at thermometry may be discussed, but it is wise to show the students how any kind of temperature scale can be constructed from standards such as the BP and FP of water. The necessity for an absolute temperature scale having no negative numerical values might be brought up. The students may then be given conversion formulas and problem drills.

Using the simple equation $\Delta Q = mc\Delta t$, the students may be given simple problems in heat transfer. The formula may be used to show how the definitions for the calorie and BTU were derived. It is important that heat transfer problems dealing with the change of state, in which the term Δt has no physical significance, be fully explained. Cooling curves, especially those showing heat effects at the freezing point and boiling point, should be covered in detail so that the change of state may be understood.

During the discussion of heat capacities, the abnormally high heat capacity of water may be discussed and also the beneficial effect of this as a heat exchanger in biochemical processes. If the concept of hydrogen bonding has been covered previously, it may be covered again with reference to these heat capacities.

As a laboratory experiment for this unit, a cooling curve may be easily constructed by melting a test tube half-full of acetamide, suspending a thermometer in the melt, and taking readings every minute. A graph of time vs. temperature should yield an excellent cooling curve, although some curves may show a slight depression in the fusion area due to supercooling.

OBEJCTIVES

1. The student should be able to demonstrate by examples the difference between temperature and heat content.
2. The student should be able to explain the three most common temperature scales which exist.

3. The student should be able to write the following temperature conversion formulas:
 - a. Fahrenheit to Centigrade
 - b. Centigrade to Fahrenheit
 - c. Centigrade to Kelvin
 - d. Kelvin to Centigrade
4. The student should be able to solve problems involving each of the above conversion formulas.
5. The student should be able to define each of the following terms:
 - a. Heat of fusion
 - b. Heat of vaporization
 - c. Specific heat
6. The student should know and be able to define each of the following heat content units:
 - a. Calorie
 - b. Btu
7. The student should be able to make the following heat unit conversions:
 - a. Calories to Btu's
 - b. Btu's to calories
8. The student should be able to explain the difference between a calorie and a Calorie (kilocalorie) and be able to convert from one unit to the other.
9. The student should be able to solve the following types of heat problems:
 - a. Amount of heat absorbed or released in changing liquid water at any temperature to liquid water at any other temperature.
 - b. Amount of heat required to change ice at any temperature to liquid water at any temperature.
 - c. Amount of heat required to change liquid water at any temperature to steam at any temperature.
10. The student should be able to interpret the formula $\Delta Q = mc\Delta t$ and then use it in solving heat problems mentioned in item nine above.
11. The student should be able to interpret at least two cooling and heating curves.

HEAT

CONTENT	LEARNING EXPERIENCES
A. Comparison between Heat and Temperature 1. Definition of temperature 2. Definition of heat B. Temperature Scales 1. Fahrenheit 2. Centigrade and Celsius 3. Absolute temperature scales a. Kelvin b. Rankin (optional) 4. Conversion formulas $^{\circ}\text{C} = 5/9(^{\circ}\text{F} - 32^{\circ})$ $^{\circ}\text{F} = (9/5^{\circ}\text{C}) + 32^{\circ}$ $^{\circ}\text{K} = ^{\circ}\text{C} + 273^{\circ}$	A Pour boiling water into 100 ml beaker and into 400 ml beaker. Record the temperature of each fifteen minutes later and compare. Discuss the difference between heat and temperature. B-3 Show how any absolute temperature scale can be derived from absolute zero. $^{\circ}\text{R} = ^{\circ}\text{F} + 460^{\circ}$ $^{\circ}\text{K} = ^{\circ}\text{C} + 273^{\circ}$ B-4 Assign problems in temperature scale conversions. B Assign class report on temperature measuring devices--thermocouples; optical pyrometers, gas thermometers, etc.
TEXT REFERENCES	AUDIO-VISUAL AIDS
Baker, Bradbury, et al., pp. 119-120; 159-161 Metcalf, Williams, Castka, pp. 14-15 Smoot, Price, Barrett, pp. 12-13	<u>McGraw-Hill Film</u> TEMPERATURE AND MATTER, #612015 Properties of matter under varying temperature conditions are discussed. The film not only deals with transformation of physical state, but also with unusual properties at low temperatures, such as electrical conductivity of metals and paramagnetism of liquid oxygen. Also includes viscosity changes of gases and liquids as temperature increases. (15 min. Color)

HEAT

CONTENT	LEARNING EXPERIENCES
C. Simple Changes in Heat Content 1. Principles of heat exchange 2. Thermal equilibrium 3. $\Delta Q = mc \Delta t$ 4. Definition of calorie and BTU D. Heat and Change of State 1. Heat of fusion 2. Heat of vaporization 3. Average specific heat 4. Cooling and heating curves	C Assign problems dealing with simple changes in heat content. Discuss difference between calorie and kilocalorie. D Laboratory Experiment: Construction of Cooling Curve (Melt acetamide; insert thermometer and record temperature every minute. Graph time vs. temperature.)
TEXT REFERENCES	AUDIO-VISUAL AIDS
Baker, Bradbury, et al., pp. 160-161 Metcalf, Williams, Castka, p. 15 Smoot, Price, Barrett, pp. 90-93	68

GASES

THE APPROACH

The unit on "Heat" should be covered immediately prior the unit on gases for maximum understanding. This will allow the concepts of temperature, heat, and energy to remain fresh in the students' minds when they are discussed under the gas laws.

The gas law's unit can be introduced by comparing the three states of matter according to shape, volume, compressibility, and diffusion. The tendency toward increased disorder in the gaseous state can be demonstrated here and conditions necessary for measurement may follow. It might be well to review the concepts of pressure as a force per unit area and then as the product of the height and density of a liquid in order to clarify the use of mm of Hg (Torr) as a pressure unit. The concept of STP can be used to emphasize the use of a standard in any kind of scientific work.

The postulates of the Kinetic Theory of gases should be discussed by using examples for each. The analogy of Elbert C. Weaver, comparing gas molecules in a container to one jar full of insane bees, will not only introduce humor into the classroom but will also demonstrate the motions of gas molecules well. Energy distribution of gaseous molecules can be compared to IQ distribution, height or weight distribution, etc. A good example is that of a clothing store--a man wearing a size 54 does not have the same selection of color and style as one wearing the more common size 40 because of the size distribution in the general population.

Following introduction of the postulates of the Kinetic Theory, ideal and non-ideal gases may be discussed. Reasons why CO_2 and NH_3 are less ideal than He may be shown in conjunction with three types of molecular motion.

At this point the students should be ready for the introduction of law problems. Students of good math background have little trouble deriving Charles' Law from laboratory data in the form $V = KT$ or $V \propto T$. Boyle's Law may be demonstrated by the instructor since the usual apparatus is expensive. The IPS method, using a dowel inside a test tube and bricks, should enable all students "to discover" Boyle's Law.

In working gas law problems, many teachers prefer to utilize what is sometimes referred to as the "Intuitive" approach, wherein no formula is given, but a pressure fraction and a temperature fraction are multiplied by the volume to secure the new volume. Teachers who like this method claim it forces the student to reason more and thus gain a better understanding of Boyle's and Charles' Laws. Although the claim may have merit, on many occasions both methods of approach should be demonstrated. A knowledge of both assures at least a better understanding of the tools necessary to attack the problems.

Consider the following problem:

EXAMPLE: What is the volume of a gas at STP if its volume at 800 mm pressure and 17° C is 25 ml?

Using Formula, Placing
Emphasis on Unit Analysis

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

$$\frac{800 \text{ mm (25 ml)}}{290^\circ \text{ K}} = \frac{760 \text{ mm } V_2}{273^\circ \text{ K}}$$

$$\frac{(800 \text{ mm})}{(760 \text{ mm})} \times \frac{(273^\circ \text{ K})}{(290^\circ \text{ K})} \times (25 \text{ ml}) = V_2$$

$$(1.05) (.975) (25 \text{ ml}) = V_2$$

$$25.6 \text{ ml} = V_2$$

Intuitive Approach
Including Units Analysis

$$V = \frac{273^\circ \text{ K}}{290^\circ \text{ K}} \times \frac{(800 \text{ mm})}{(760 \text{ mm})} \times 25 \text{ ml}$$

$$V = (.975) (1.05) 25 \text{ ml}$$

$$V = 25.6 \text{ ml}$$

The formula method coincides with the "intuitive method", and all units cancel nicely. For the slide rule, either set-up is effective.

Graham's Law of Diffusion may be inserted before Charles', Boyle's, and Gay-Lussac's Laws as part of the Kinetic Theory of gases. The old reliable diffusion experiment from the MCA series using combustion tubes, $\text{NH}_3(g)$, and $\text{HCl}(g)$ is an excellent one which may be utilized in a variety of ways to determine densities or molecular weights or to compare relative diffusion rates of gaseous molecules.

Dalton's Law of Partial Pressures, corrections for vapor pressure of water, and corrections for differences in levels of hydrostatic pressure may be omitted with less proficient groups.

The Ideal Gas Law may also be omitted with these groups, but it should be covered with faster groups. The Victor Meyer experiment for determination of gas densities may be performed with relatively simple equipment to illustrate how the Ideal Gas Law can be used to find molecular weights. Write to SMIP for #CC-219.

OBJECTIVES

1. The student should be able to differentiate the solid, liquid, and gaseous states of matter, and give a set of characteristics peculiar to each.
2. The student should recall that a knowledge of PV is necessary for the measurement of gases.
3. The student should be able to recall five of the six postulates of the Kinetic Molecular Theory.
4. The student should be able to explain a plot of energy distribution of gas molecules versus numbers of molecules in energy states.
5. Given the molecular structure of a compound, the student should be able to predict which intramolecular motions are present.
6. Given the necessary information, the student should be able to solve the following types of problems:
 - a. Boyle's Law
 - b. Charles' Law
 - c. Gay-Lussac's Law
 - d. The General Gas Law
 - e. Vapor Pressure Corrections
 - f. Corrections for Differences in Levels (optional)
 - g. The Ideal Gas Law (optional)

GASES

CONTENT	LEARNING EXPERIENCES
A. Defintion (Gas has no definite volume or shape). B. Relation of Gases to Liquids and Solids 1. More molecular disorder in gases 2. Conditions necessary for description of equation of state a. Pressure b. Volume c. Temperature d. Mass e. Standard conditions C. Kinetic Molecular Theory 1. Postulates of the Kinetic Theory	B Disscuss definitions of pressure and formulas from physics. $P = \frac{F}{A} \quad \text{and} \quad P = hD$ C Compare gas molecules to Weaver's Analogy - jar full of insane bees. If CENCO Kinetic Theory apparatus is available, demonstrate. Discuss volume of 1 g of water; as liquid 1 g H_2O = 1 ml at 4° C; as gas 1 g H_2O = 1700 ml at 100 ° C. Show elastic collision with rail and marbles.
TEXT REFERENCES	AUDIO-VISUAL AIDS
Baker, Bradbury, et al., pp. 109-113 Metcalf, Williams, Castka, pp. 125-130 Smoot, Price, Barrett, pp. 209-214	<u>Chem Study Film</u> MOLECULAR MOTIONS, #4115 Many properties of matter such as fluidity, vaporization and rates of chemical reactions, indicate that molecular motion must be occurring, and that the freedom of motion increases in going from the solid to the liquid to the gaseous state. The concepts of translational, rotational, and vibrational molecular motions allow the interpretation of the observed properties. The use of animation and dynamic models makes clear how the observed properties depend on the types of motion occurring at the molecular level. (13 min. - Color) <u>McGraw-Hill Filmstrip</u> KINETIC MOLECULAR THEORY

GASES

CONTENT	LEARNING EXPERIENCES
C (Continued) 2. Van der Waals forces and KMT 3. Energy distribution of gas molecules a. Gaussian distribution b. Average energy 4. Ideal gases a. Types of molecular motion i. Translatory ii. Vibratory iii. Rotatory b. Ideality in gases i. PV product (Boyle's Law) = K at const. T ii. Obey postulates of KMT iii. Complex molecules and ideality 5. KMT and vapor pressure 6. Graham's Law of Diffusion $\frac{R_1}{R_2} = \frac{\sqrt{MW_2}}{\sqrt{MW_1}}$	C-2 Use graphs of height distribution, weight distribution, etc. in class. Discuss clothing store order. Compare to electron probability distribution. C-3 Demonstration of random distribution using the overhead projector and probability distribution apparatus from 3M. C-4 Use ball and spring models of ethane, ethylene, and acetylene to illustrate types of molecular motion. For humor, compare to modern dances. C-5 Discuss pressure cooker: discuss problems in cooking at high altitudes. C-6 Ringing bell with hydrogen
TEXT REFERENCES	AUDIO-VISUAL AIDS
Baker, Bradbury, et al., p. 113; 114 Metcalf, Williams, Castka, pp. 125-130 Smoot, Price, Barrett, pp. 210-217 223-225 268-270 <u>The Gas Laws</u>, Sackheim, George, Harper & Row. (Programmed text) <u>Gases: Kinetic Theory of Gases</u> <u>Gases: Gas Laws</u>, Ford, Renee G., Holt, Rinehart & Winston. (Programmed text)	<u>Chem Study Film</u> VIBRATION OF MOLECULES, #4118 All animation. The film shows the relationship between the structure of a molecule and its vibrational motions. Water, carbon dioxide, and methane are discussed in detail. The forms of the vibrations have been accurately calculated from spectral data. All vibrations have been slowed down by a factor of 10^{14}. The effect of molecular collision, or absorption of light, on molecular vibrations is illustrated. Determination of the number of possible vibrations and the analysis of complex vibrations in terms of simple harmonic motions are explained. (12 min. - Color)

GASES

CONTENT	LEARNING EXPERIENCES
D. Gas Laws 1. Charles's Law - Need for absolute temperature scale - Derivation $V = KT$ at const. P - Usual formula $\frac{V_1}{V_2} = \frac{T_1}{T_2} \qquad \frac{V_1}{T_1} = \frac{V_2}{T_2}$ d. Failure at low temperature 2. Boyle's Law - Comparison of pressure units - Derivation $PV = K$ (const. T) - Usual formula $P_1V_1 = P_2V_2 \text{ (const. } T\text{)}$ d. Failure at high pressure P	D-1 <u>LABORATORY EXPERIMENT:</u> "Discovery" of Charles' Law Assign problems on Charles' Law. D-2 <u>LABORATORY EXPERIMENT:</u> "Discovery" of Boyle's Law. - Usual Boyle's Law apparatus as demonstration. - IPS method for Boyle's Law (SMIP #CC230) Assign problems on Boyle's Law.
TEXT REFERENCES	AUDIO-VISUAL AIDS
Baker, Bradbury, et al., pp. 116-120 Metcalf, Williams, Castka, pp. 133-136 Smoot, Price, Barrett, pp. 271-274 279-282	<u>PSSC Film</u> BEHAVIOR OF GASES Shows brownian motion and discusses Boyle's Law. (15 min. - B/W) <u>Coronet Film</u> THE LAWS OF GASES The Gas Laws of Boyle, Charles, Dalton, and Avogadro are explained and their importance to physical science is emphasized. Animation clearly illustrates relationships between pressure, volume and temperature of confined gas, the law of partial pressures, and the determining of the molecular weight of a gas. (11 min. Color \$120 - B/W \$60)

GASES

CONTENT	LEARNING EXPERIENCES
D (Continued) 3. Gay-Lussac's Law - Use in thermometry - Derivation $P = KT$ (const. V) - Usual formula $\frac{P_1}{T_1} = \frac{P_2}{T_2}$ 4. Dalton's Law of partial pressures - Definition of partial pressure - Usual formula $P_t = P_1 + P_2 + P_3 + \dots + P_n$ - Application to vapor pressure correction 5. Corrections for differences in levels - Collection of gases over Hg - Collection of gases over H_2O 6. The Ideal Gas Law - Review of ideality - Derivation of $PV = nRT$ - Usual formulas $PV = nRT \text{ and } PV = \frac{gRT}{M}$ - Victor Meyer method for mw determination	D-3 Discuss Beckmann constant-volume gas thermometer. D-4 Discuss significance of KMT on Dalton's Law. Assign problems involving vapor pressure correction for gases collected over water. D-6 Derive Ideal Gas Law. <u>LABORATORY EXPERIMENT:</u> Determination of molecular weight of $HC_2H_3O_2$ vapor by Victor Meyer Method. (SMIP # CC219) Assign problems on Ideal Gas Law; stress mw determination.
TEXT REFERENCES	AUDIO-VISUAL AIDS
Baker, Bradbury, et al., pp. 120-122 Metcalf, Williams, Castka, pp. 131-132, 138; 152-155 Smoot, Price, Barrett, pp. 277-278	<u>Chem Study Film</u> GAS PRESSURE AND MOLECULAR COLLISIONS The film explores the relationship between gaseous pressure and molecular collisions. The effects of varying the number of molecules per unit of volume and of varying the temperature are studied. The experimental study of the relative rates of effusion of hydrogen, oxygen, carbon dioxide and sulfur hexafluoride leads to the quantitative relationship between molecular weight, molecular velocity and absolute temperature. Mechanical models illustrate the experimental observations.

WATER

THE APPROACH

The importance of water in our daily lives from almost every standpoint dictates covering this topic as an ordinary unit. Many do not place water in its proper perspective as a standard and the basis for many comparisons of matter. Many students do not relate the 1.0 g/cc density of water, the 1.0 Cal/g. deg C heat capacity, and the mp and bp of water as bases for standardization and as being inter-related.

The physical properties of water may be demonstrated quickly by means of discovery-type demonstrations; the chemical properties may also be demonstrated in the same way. It is suggested that the chemical properties of water be stressed through general word equations.

If not covered earlier, the nomenclature of hydrates may be appropriately discussed here.

A most important concept in the lesson is hydrogen bonding in water. This concept should be stressed; most students will be shocked to learn that water would be a gas at room temperature if no hydrogen bonding existed. Other abnormal properties of water should be explained and compared. Using graphs of similar compounds and the periodic table, the unique properties of water can be easily shown.

If the properties of heavy water (deuterium oxide) were not covered earlier in the section on isotopes, the properties of heavy water may be compared using the SMIP filmstrip, "Properties of Isotopes".

OBJECTIVES

1. The student should be able to state four ways in which water is used as a standard in scientific measurement.
2. The student should be able to state at least four methods used for purifying water.
3. The student should be able to state at least six physical properties of water.
4. The student should be able to state at least four chemical properties of water by writing the product(s) of the following reactions:
 - a. Reaction between an active metal and water
 - b. Reaction between a metallic oxide and water
 - c. Reaction between a non-metal and water
 - d. Reaction between a non-metallic oxide and water

5. The student should be able to explain the meaning of each of the following terms:
 - a. Efflorescence
 - b. Deliquescence
 - c. Desiccation
6. The student should be able to explain the process of hydrate formation in a given hydrate.
7. The student should be able to show how the following terms apply to the structure of a water molecule:
 - a. Covalency
 - b. Nonlinearity
 - c. Polarity
8. The student should be able to explain the basic difference between hydrogen bonding and ordinary covalent bonding, using a combination of water molecules as an example of hydrogen bonding and an individual water molecule as an example of covalent bonding.
9. The student should be able to specifically point out how hydrogen bonding in water is responsible for its unique character among liquids of being most dense at a temperature other than its freezing point.
10. The student should be able to explain the difference in density between H_2O and ice from the standpoint of hydrogen bonding.
11. The student should be able to give the structural arrangement of atoms in D_2O and be able to contrast at least two ways in which the properties of D_2O are different from its isotopic counterpart, H_2O .

WATER

CONTENT	LEARNING EXPERIENCES
A. Introduction 1. Need for H₂O in life 2. Need for standard a. Specific gravity b. Weight c. Volume d. Thermometer -temperature scales e. Heat i. Definition of calorie ii. Definition of B. T. U. 3. Low cost 4. Easily purified 5. Ready availability B. Physical Properties 1. Density 2. Color 3. Odor 4. Taste	A a. Student experiment on purification of H₂O by distillation. May also compare distillation and ion exchange resins as methods of purification. b. Demonstrate comparative densities by contrasting H₂O floating on CCl₄ with H₂O and alcohol. Expand to use of H₂O as standard for specific gravity. c. Recall definition of gram and liter and fixed points on thermometer scales. B Compare properties of pure water (distilled) and tap water.
TEXT REFERENCES	AUDIO-VISUAL AIDS
Baker, Bradbury, et al., pp. 145-150; 159-163 Metcalfe, Williams, Castka, p. 172 Smoot, Price, Barrett	<u>Northeastern Pennsylvania Audio-Visual Center</u> THE WORLD OF H₂O <u>Welch Filmstrip, Group I</u> HYDROGEN, OXYGEN AND H₂O, # 3992R

WATER

CONTENT	LEARNING EXPERIENCES
C. Chemical Properties 1. Reaction with active metals 2. Reaction with metallic oxides 3. Reaction with non-metals 4. Reaction with non-metallic oxides	C Demonstrate reactions of Na, CaO, and CO₂ with H₂O, testing resulting solutions with litmus. Discuss purification of H₂O by chlorination. Assign class report on importance of water in our daily lives--biologically, chemically, industrially. An alternative assignment could deal with water pollution.
TEXT REFERENCES	AUDIO-VISUAL AIDS
Baker, Bradbury, et al., p. 164 Metcalf, Williams, Castka, p. 175 Smoot, Price, Barrett	<u>Coronet Films</u> CHEMICAL PROPERTIES OF WATER The chemical properties of water, our most common chemical compound are explained in terms of the polar nature of the water molecule. The broad range of chemical reactions of water with metals, metal oxides, non-metal oxides, and salts is demonstrated. We see how water in its important role of a solvent, promotes other chemical reactions. (Color - 13 1/2 min.)

WATER

CONTENT	LEARNING EXPERIENCES
C (Continued) 5. Formation of hydrates - Efflorescence - Deliquescence - Desiccation D. Structure - Covalency - Nonlinearity - Polarity	C-5 a. Demonstrate changes of $\text{Na}_2\text{CO}_3 \cdot 10 \text{H}_2\text{O}$ and CaCl_2 granules upon exposure of the atmosphere. b. Demonstrate change in $\text{CuSO}_4 \cdot 5 \text{H}_2\text{O}$ upon heating and anhydrous CuSO_4 upon addition of H_2O. Use transparent vycor crucible. c. Student experiment on determination of the percent H_2O in a hydrate. D-1 Compare heating of H_2O and HgO to emphasize strength of covalent bonds in H_2O. Compare properties of water with H_2S and other congeners.
TEXT REFERENCES	AUDIO-VISUAL AIDS
Baker, Bradbuty, et al. Metcalf, Williams, Castka, p. 173 Smoot, Price, Barrett	<u>Northeastern Pennsylvania Audio-Visual Center</u> THE STRUCTURE OF H_2O, # 1166, (14 min. - Color) <u>FOM Filmstrip</u> WATER --A MOST UNUSUAL SUBSTANCE, # 634 Water's polar nature and consequent hydrogen bonding provide explanations for its unusual physical properties such as high freezing and boiling points, optimal density at 4°C, and high specific heat. Its chemical properties are illustrated by many experiments. Some aspects of solution bonding and its chemical properties are suitable for general science. For school systems in which the study of water has become part of general science programs, the related frames can serve as valuable reviews and summaries of concepts and experiments.

WATER

CONTENT	LEARNING EXPERIENCES
E. Description of Hydrogen Bonding in Terms of: 1. Electrostatic attractions 2. Effect on boiling point, melting point, and heats of vaporization and fusion 3. Effect on comparative densities of H₂O and ice F. D₂O 1. Occurrence 2. Separation from H₂O 3. Uses	D-2 Stress deviation of H₂O from general trend in properties of hydrides of Group VI with increasing molecular weight. D-3 Use density graphs or tables to illustrate unusual behavior of H₂O. (Modern Physics, 1964, p. 237) F Student report on isotope effects
TEXT REFERENCES	AUDIO-VISUAL AIDS
Baker, Bradbury, et al., pp. 155; 165-167 Metcalf, Williams, Castka, pp. 174; 173 Smoot, Price, Barrett, pp. 234-237	<u>Film Associates</u> PHYSICS AND CHEMISTRY OF WATER Water is one of the most extraordinary substances known. Life on earth depends upon some of its unusual properties--its slow rate of evaporation, its high surface tension, its power of solution, and the fact that it is heavier in the liquid than the solid form. The water molecule is made of two atoms of hydrogen and one of oxygen, held together by electrical forces. One end of the water molecule is positively charged and the other end is negatively charged. The positive end of one water molecule can be attracted and attached to the negative end of another, forming a hydrogen bond. These bonds are continually being made and broken at room temperature, resulting in fluidity. Hydrogen bonds account for most of water's unusual properties. (21 min. - Color \$ 12.50)

SOLUTIONS

THE APPROACH

The unit may be opened with reference to a simple solution of sugar and water. A volunteer in the class may be asked to taste what apparently is a cup of drinking water; detection of the sugar solute from the sweet taste can be used to introduce the concepts of solute and solvent. A chart of the nine different possible solution types can be constructed, using examples from the students' own experience. The teacher may augment the chart if difficult examples cannot be readily recalled. The concept of ideal and non-ideal solutions with regard to solute/solvent interactions may be demonstrated by showing that $50 \text{ ml EtOH} + 50 \text{ ml H}_2\text{O} = 95 \text{ ml solution}$ instead of the expected 100 ml.

The concept of ionization may be delayed until the unit on ionization, if it is presented later when solutes are discussed from a simple ionization standpoint. The solvation process can be handled with an inquiry lecture-demonstration approach on the factors which affect solubility.

A few solubility curves should be covered in the lab before discussing them in the classroom, thus emphasizing the discovery process. With classes of average ability, salts such as KCl , $\text{Pb}(\text{NO}_3)_2$, KNO_3 , and $\text{K}_2\text{Cr}_2\text{O}_7$ may be used in the lab experiment. With classes of greater ability, while more unusual solubility curves (such as $\text{Ce}_2(\text{SO}_4)_3$ and NaCl) may be assigned. The relation of the shape of the curve to the heat of solution should be covered in the class lecture.

The effect of solutes on the boiling point and freezing point of a solvent may now be illustrated. If vapor pressure curves of the pure solvent and solvent + 1 mole of solute are compared, Raoult's Law and problems involving boiling point and fp may be covered with an opportunity for greater student understanding. The method of solving bp elevation and fp depression problems is more meaningful if the concept of molal solutions is introduced first. If the student understands the molal concentration of a solution, then the simple relationship $\Delta T = K_m$ can be used to solve fp and bp problems.

After the student comprehends the solvation process, water of hydration may be covered. Determination of the water of hydration of BaCl_2 and computation of its formula is a good illustrative experiment because it demonstrates both the concept of water of hydration and the concept of determination of simple formula.

OBJECTIVES

1. The student should be able to name the components of a solution.
2. The student should be able to differentiate between ideal and non-ideal solutions.
3. The student should be able to explain the conditions that lead to solution equilibrium.
4. The student should be able to recall four factors that affect the solute dissolution rate.
5. The student should be able to distinguish the difference between an unsaturated solution, a saturated solution, and a supersaturated solution.
6. The student should be able to construct a solubility curve from experimental evidence.
7. The student should be able to interpret a given solubility curve.
8. The student should be able to explain the difference between a positive and a negative heat of solution.
9. The student should be able to state LeChatelier's Principle, and fully explain its implications in the process of solution formation.
10. The student should be able to explain the process of crystal formation and also the conditions which are necessary for crystallization to occur.
11. The student should be able to explain four ways of expressing solution concentration.
12. The student should be able to solve the following types of problems with non-ionic solutes:
 - a. Freezing-point depression
 - b. Boiling-point elevation
 - c. Calculation of molecular weights from freezing-point depression and boiling-point elevation information

SOLUTIONS

CONTENT	LEARNING EXPERIENCES
A. Definition of a Solution 1. Solute 2. Solvent 3. Non-electrolytes	A Exhibit NaCl, EtOH, H ₂ O. Show how they can be mixed to form solutions. Show how NaCl solution can be evaporated to dryness on watch glass to secure solute.
B. Types of Solutions 1. Ideal 2. Non-ideal	Perform Alyea demonstration 5-9. Mix 50 ml EtOH and 50 ml H ₂ O in volumetric flask: show that 95 ml solution is formed.
TEXT REFERENCES	AUDIO-VISUAL AIDS
Baker, Bradbury, et al., p. 168 Metcalfe, Williams, Castka, p. 183 Smoot, Price, Barrett	<u>McGraw-Hill</u> LIQUIDS IN SOLUTION A study of the theoretical basis for understanding why some liquids are miscible and some are immiscible, and how this may explain the formation of emulsions as well as the process of liquid extraction. [11 min. - Color (Code 612014) - B/W (Code 612002)] <u>FOM Filmstrip</u> THE SOLUTION PROCESS, #661 Students learn the difference between solutions as homogeneous mixtures, and coarse suspensions and colloidal dispersions as heterogeneous mixtures; the role of water as a common solvent is discussed. The interactions involved in the dissolving process and the accompanying energy changes are explained, as are the effects of changing conditions on the solubility of a variety of solutes.

SOLUTIONS

CONTENT	LEARNING EXPERIENCES
C. Equilibrium in Solutions 1. Dynamic nature of solubility equilibrium 2. Saturated solutions 3. Mechanism of solution process 4. Factors affecting solubility a. Temperature b. Pressure (Henry's Law) c. Nature of solvent and solute d. Surface area of solute e. Stirring	C-1 Dynamic nature of solubility can be shown with radioactive I^{131} and AgI ppt. or similar systems. See SMIP CC-220. C-3 Demonstrate factors affecting solubility a. Add KNO_3 to cold water and show how some solute remains on bottom; heat. b. Alyea demonstration 5-12 c. Add CCl_4 to H_2O - doesn't mix. Add CCl_4 to hexane - mixes. Add CH_3OH to CCl_4 and H_2O, shake. Add naphthalene to CCl_4 and then H_2O. d. Dissolve sugar cube, then grind up another sugar cube and dissolve. Show how stirring increases rate of solution.
TEXT REFERENCES	AUDIO-VISUAL AIDS
Baker, Bradbuty, et al., pp.171-179 Metcalf, Williams, Castka, pp.184-187 194-196 Smoot, Price, Barrett <u>Dissolving</u>, Nuffield #CBB04335 Penguin Books	<u>Cornet Films</u> SOLUTIONS The more important characteristics of solutions are presented by means of laboratory demonstrations. Both ionic and non-ionic solutions are defined, and the common types visualized. We learn of the factors which influence solubility such as temperature, pressure, and composition. In conclusion, see some of the practical applications of solutions in the chemistry of everyday life. (16 min. - Color - B/W) EBF PROPERTIES OF SOLUTIONS The natures of solutes and solvents in the preparation of solutions are demonstrated. Solutions of slightly soluble substances, factors influencing the solution of one liquid in another, and the effect of temperature on the solubility of a solute are taken up. Supersaturation and the principle of extraction are discussed and demonstrated. (28 min. - Color - B/W)

SOLUTIONS

CONTENT	LEARNING EXPERIENCES
D. Solubility Curves 1. Definition 2. Types of solutions a. Dilute b. Concentrated c. Saturated d. Supersaturated 3. Heat of solution a. Positive b. Negative c. Meaning of convention d. LeChatelier's Principle 4. Crystallization	D <u>Laboratory Experiment:</u> Determining a solubility curve. Assign KNO_3 - $\text{Pb}(\text{NO}_3)_2$ - NaCl - $\text{K}_2\text{Cr}_2\text{O}_7$ and $\text{Ce}_2(\text{SO}_4)_3$. All have a positive slope except $\text{Ce}_2(\text{SO}_4)_3$: heat of solution will explain why. NaCl exhibits a lower slope than others.
TEXT REFERENCES	AUDIO-VISUAL AIDS
Baker, Bradbury, et al., pp. 170; 179 Metcalf, Williams, Castka, pp. 188-192 Smoot, Price, Barrett	<u>McGraw-Hill</u> DYNAMICS OF SOLUTION The difference between saturated, unsaturated and supersaturated solutions are demonstrated. Characteristics of solutions, proof that solutions are non-static, and the nature of dynamic equilibrium are clearly presented. [14 1/4 min. - Color (Code 612017) B/W (Code 612005)] <u>Cornet Film</u> HEAT OF SOLUTION The heat evolved when a known amount of zinc sulfate is dissolved in water is measured and the molar heat of solution of zinc sulfate is calculated. (16 min. - Color - B/W)

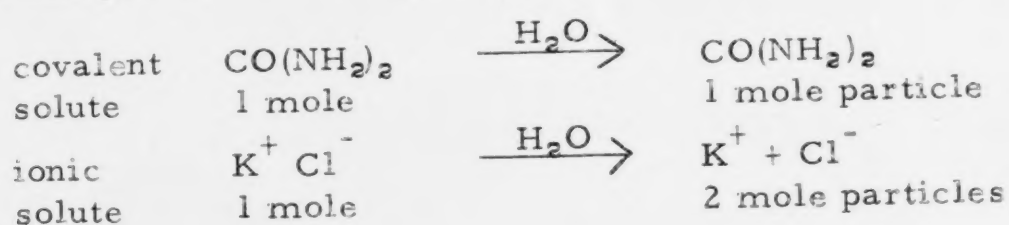
SOLUTIONS

CONTENT	LEARNING EXPERIENCES
E. Concentration of Solutions 1. Percentage concentration a. Weight-weight b. Weight-volume c. Volume-Volume 2. Molality a. Definition b. Use of molality in freezing-point depression c. Use of molality in boiling-point elevation d. Use of molality in calculation of molecular weights 3. Molarity a. Definition b. Using molar solutions	E-1 Comment on different methods of expressing percent concentration. Comment on use of proof system in alcoholic solutions. E-2 Discuss anti-freeze and FP depression; which makes better anti-freeze, low or high MW? Assign problems on FP depression and BP elevation. E-2(d) <u>Laboratory Experiment:</u> "Finding Molecular Weights" (MCA type) Find molecular weights from FP depression of urea, acetamide, chloral hydrate, sucrose. E-3 Assign problems using molar solutions.
TEXT REFERENCES	AUDIO-VISUAL AIDS
Baker, Bradbury, et al., p. 179 Metcalfe, Williams, Castka, pp. 193; 197-199 Smoot, Price, Barrett	<u>McGraw-Hill, Sec 5</u> SOLUTIONS OF DEFINITE CONCENTRATION, # 019210 <u>Welch Scientific Company, Group I</u> SOLUTIONS, SUSPENSIONS, AND COLLOIDS, #3992R <u>Coronet Film</u> MOLECULAR WEIGHT OF SOLUTES Methods of the determination of molecular weights of non-volatile solutes are discussed. Two such determinations are made using an organic solvent and water, respectively. The conventional calculations are shown and discussed. (8 min. - Color - B/W)

IONIZATION

THE APPROACH

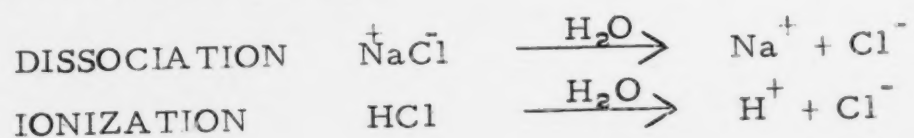
The background for the unit on ionization may be laid late in the unit on solutions. If the MCA experiment, "Finding Molecular Weights", is used, one group should be assigned sodium chloride or some other similar monovalent salt as one of the compounds. Since ionic compounds yield more particles in solution than predicted from theory...



...the freezing point should be greater per mole of solute for ionic compounds than for covalent compounds. The freezing point of a one molal NaCl solution is -3.6°C instead of the predicted -1.86°C . This apparent inconsistency gives the basis for the assumption that more than one mole of particles is present per 1000 g of solvent, as defined in a one molal solution. Various solutes in water can then be tested with conductivity apparatus. The types of solutes - electrolytes and non-electrolytes - can be explored. The degree of ionization may be discussed either by making a rough observation of the lighting intensity or by connecting an electrical indicator to the circuit.

The MCA experiment or other similar experiments, "Conductivity of Solutions", may be investigated either as a demonstration or a lab experiment. Dissociation into ions, when salts are melted, can be demonstrated with a low-melting point salt such as potassium iodide.

The difference between dissociation and ionization should be stressed. By means of elementary ionic equations, one showing the ionization of acetic or hydrochloric acid and another showing the dissociation of a salt, this concept can be understood more easily.



Once the mechanism of ionization has been grasped by the students, practice in writing ionic equations should not be delayed. Writing ionic equations encompasses a number of earlier concepts, such as equations, ionic charges, and combining numbers, and lays a foundation for later work in acids, bases, and equilibria.

OBJECTIVES

1. The student should be able to distinguish between an atom and an ion.
2. The student should be able to differentiate between an electrolyte and a non-electrolyte in terms of properties and structure.
3. The student should be able to exhibit understanding of ionic dissociation by writing complete and simple equations.
4. The student should be able to distinguish between non-ionizable and ionizable substances from a given list of compounds, with a least 80% accuracy.
5. The student should be able to show how the autoionization of H_2O helps explain the formation of hydronium ion and hydroxide ion in solution.
6. The student should be able to show the effect on the freezing point and boiling point of a solution of electrolytes and non-electrolytes to a solvent.
7. The student should be able to distinguish between a strong electrolyte and a weak electrolyte from the difference in their degrees of ionization.
8. The student should be able to show the effect a change in solution concentration has on the degree of ionization.
9. The student should be able to explain the change in electrolyte behavior with the addition of any one of a number of chemical compounds.
10. The student should be able to explain fully what part the concept of ionization plays in electrolysis (a specific compound should be chosen for this explanation).

IONIZATION

CONTENT	LEARNING EXPERIENCES
A. Types of Solutes 1. Non-electrolytes 2. Electrolytes B. Theory of Ionization 1. Structure of electrolytes 2. Dissociation 3. Ionization	A Demonstrate Alyea demonstration 6-1 or 6-2; show conductivity of salts vs. non-conductivity of compounds such as sugar, ethyl alcohol, methyl alcohol, urea - aq. etc. Alyea demonstration 6-7. Add Mg to dilute HCl and toluene; HCl reacts, toluene does not. B Use transparencies to show differences between atoms and ions, mechanism of ionization.
TEXT REFERENCES	AUDIO-VISUAL AIDS
Baker, Bradbury, et al., pp. 237-241 Metcalf, Williams, Castka, pp. 202-204 Smoot, Price, Barrett	<u>Coronet Film</u> IONIZATION The theory of ionization is treated with emphasis on its present important status in chemistry. The theory is defined, and we see the identifying characteristics of ions and where they occur. Electrolytes, dissociation, solvents, electrovalent compounds, and covalent compounds are shown by animation and laboratory demonstration. In conclusion reference is made to important applications of ionization in industrial chemistry. (18 1/2 min. - Color \$210 - B/W \$105)

IONIZATION

CONTENT	LEARNING EXPERIENCES
C. Electrolyte Behavior 1. Electrolysis of water 2. Effect on FP and BP 3. Degree of ionization a. Strong electrolytes b. Weak electrolytes c. Concentration effect 4. Precipitation D. Writing Ionic Equations	C Demonstrate (Alyea 1-4) electrolysis of water. Find FP of 1 m NaCl and 1 m urea solutions in ice bath. Discuss discrepancies in FP depression. Show apparent degree of ionization with Alyea 6-13. D Practice writing ionic equations.
TEXT REFERENCES	AUDIO-VISUAL AIDS
Baker, Bradbury, et al., pp. 241-244 Metcalf, Williams, Castka, pp. 204-210 Smoot, Price, Barrett	<u>Coronet Film</u> PRINCIPLES OF IONIZATION The effects of different types of solutes on the boiling point of water are shown; next the electrical conductivity or lack of conductivity of these solutions is demonstrated, and the results are correlated. The importance of the choice of solvent is shown. Ion migration during electrolysis is demonstrated, using a solution containing a mixture of copper sulfate and potassium dichromate. (13 1/2 min. - Color - B/W)

ACIDS, BASES, SALTS

THE APPROACH

The unit on acids, bases, and salts can be covered most comfortably if it follows the unit on ionization or some discussion of ionization. A number of concepts involved with acid-base theory should be discussed if students are expected to receive an understanding of this most important subject. These are:

1. Beginning simple reaction: acid + base $\rightarrow$ salt + water
2. Arrhenius Theory - proton donor (acid) and hydroxyl ion donor (base). First definition.
3. Bronsted-Lowry Theory - proton donor (acid) and proton acceptor (base).
4. Lewis Theory - electron pair acceptor (acid) and electron pair donor (base).

Similarities of all three theories should be demonstrated as each is introduced.

Simple demonstrations of the properties of acids, bases, and salts may be used to demonstrate the elementary concept: acid + base $\rightarrow$ salt + water, utilizing taste (dilute acetic acid) and touch (dilute sodium hydroxide). Combining equal proportions of each destroys both properties.

Following the unit on ionization, the unit on acids, bases, and salts may be opened with a few simple demonstrations interspersed with discussion and notes on the demonstrations. One effective demonstration involves either the class clown or class tough guy. A bottle of glacial acetic acid - with the POISON declaration facing the class - is placed on the demonstration table beforehand. A dilute solution of acetic acid, about 3-5%, is made from the glacial acetic acid, and a volunteer is requested from the two individuals mentioned above. If they or others do not taste the vinegar solution, the teacher may do so. This simple demonstration unlocks the door for other demonstrations exploring the properties of acids and bases. The general word equation for the neutralization of acids and bases should be given at this time:

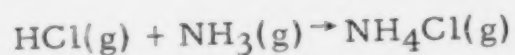


Following a demonstration of simple acid-base reactions, the Arrhenius Theory may be introduced pointing out that the discovery of the theory was closely allied to the discovery of the theory of electrolytes' dissociation in water. Sometime during the discussion the color-coding used by the chemical industry should be explained to the students.

red - nitric acid
 blue - hydrochloric acid
 yellow - sulfuric acid
 brown - acetic acid
 green - ammonium hydroxide
 white/black - all others

If possible, the laboratory should adopt this color-coding for the negative radicals of stock salt solutions to prevent mixing of chemicals accidentally.

The Bronsted-Lowry and Lewis acid-base theories may be introduced by another simple demonstration. A small beaker of conc. HCl is placed next to a small beaker of conc. NH_4OH ; a white mist of NH_4Cl appears rapidly. The students are reminded that this is an acid-base neutralization because a salt is formed.



Yet no water is present! The Bronsted-Lowry theory can now be introduced at the proper moment, along with the all-encompassing Lewis Theory.

It is imperative that the mechanism for hydration of acids involving the formation of the hydronium ion be covered in detail, using Lewis dot formulas and ionic equations. This section usually presents some difficulty to many students; and therefore, it may be necessary to present the concept quite frequently.

During the mechanism lesson, the concept of amphiprotic substances such as Al(OH)_3 and H_3AlO_3 may be introduced, using the same type of formulas.

Once the student has mastered the foregoing concepts, the pH scale may then be introduced. The pH of common materials may be examined with Hydrion papers or a pH meter if one is available. Simple pH calculations can be assigned with average classes.

Salts can be covered as the last part of the unit, with general word equations for the formation of salts, followed by hydrolysis and simple mechanism of hydrolysis.

A simple titration of vinegar with a standard base solution (about 0.50 N) could end the unit.

OBJECTIVES

1. The student should state five properties of acids.
2. The student should state five properties of bases.
3. The student should be able to state the two basic postulates of the Arrhenius Acid-Base Theory.
4. The student should be able to define an acid and a base according to the Bronsted-Lowry Theory.
5. The student should be able to identify the conjugate acid and base pairs in a given Bronsted-Lowry Acid-Base reaction.

6. The student should be able to define an acid and a base according to the Lewis Acid-Base Theory.
7. The student should be able to identify the acid and the base in a given Lewis Acid-Base reaction.
8. The student should be able to illustrate H_3O^+ ion formation using an ionic equation.
9. The student should be able to define an amphiprotic substance and illustrate his understanding by showing why $\text{A}(\text{OH})_3$ is considered amphiprotic.
10. The student should be able to explain self-ionization using H_2O molecules as an example.
11. The student should be able to show how the percentage of ionization of a certain acid or a certain base enables one to predict the relative strength of that particular acid or base.
12. The student should be able to define the concept of pH as it pertains to the measurement of the acidity or basicity of solutions.
13. The student should be able to demonstrate his ability to determine the pH of at least three solutions by using indicators.
14. The student should be able to define the concept of pOH as it pertains to the measurement of the acidity or basicity of solutions.
15. The student should be able to correctly solve four out of five problems dealing with pH and/or pOH determination.
16. The student should be able to describe two processes of salt formation.
17. The student should be able to correctly predict the effect of litmus or other indicators on the following solutions:

- a. Salt of a strong acid and a strong base
- b. Salt of a weak acid and a weak base
- c. Salt of a strong acid and a weak base
- d. Salt of a weak acid and a strong base.

(The student may verify his predictions by actual experimentation).

ACIDS, BASES, SALTS

CONTENT	LEARNING EXPERIENCES
A. Properties of Acids and Bases 1. In solution 2. Taste and feel 3. Electrolytes 4. Affect indicators 5. React with active metals B. Arrhenius Acid-Base Theory 1. Acid - liberates H^+ in aqueous solution 2. Base - liberates OH^- in aqueous solution	Demonstration: A-1 Use conductivity apparatus to show formation of ions in solution. A-2 Pour 20 ml HAc in 250 ml: taste to demonstrate sour taste of acids; pour $NaOH$ in H_2O, show slippery feel of bases. A-4 Test acids and bases with indicators. A-5 Alyea 4-12, reaction of acids with metals add Ca metal to H_2O, show H_2 formed.
TEXT REFERENCES	AUDIO-VISUAL AIDS
Baker, Bradbury, et al., pp. 238-243 245-249, 251 Metcalf, Williams, Castka, pp. 217-221 224-227 Smoot, Price, Barrett, pp. 366-367 385-386	<u>Coronet Film</u> PROPERTIES OF ACIDS, BASES, AND SALTS The conventional properties of solutions of acids and bases are illustrated and tests for these types of compounds are developed. Several reactions are then considered, among them the reaction between calcium oxide and water, zinc oxide and water, and sulfur dioxide and water. The formations of salts in several different types of reactions is demonstrated. Equations for the reactions involved above are derived and discussed. (Color - 2 1/2 reels, 28 min., B/W) <u>McGraw-Hill, Set #2 (Filmstrip)</u> ACID AND BASIC SOLUTIONS, #019005

ACIDS, BASES, SALTS

CONTENT	LEARNING EXPERIENCES
C. Bronsted-Lowry Acid-Base Theory 1. Acid - proton donor 2. Base - proton acceptor 3. Conjugate acids and bases D. Lewis Acid-Base Theory 1. Acid - electron pair acceptor 2. Base - electron pair donor	Demonstrate reaction of HCl(g) and $\text{NH}_3\text{(g)}$ to form a salt from $\text{HCl(aq)} + \text{NH}_3\text{(aq)}$ as an acid-base reaction, yet no water present.
TEXT REFERENCES	AUDIO-VISUAL AIDS
Baker, Bradbury, et al., pp. 284-289 Metcalfe, Williams, Castka, pp. 213-219 Smoot, Price, Barrett, pp. 362-366	<u>Coronet Films</u> ACIDS, BASES, AND SALTS Briefly surveying the early theories of Arrhenius and later theories of Bronsted and Lowry, and Lewis, the film uses the Arrhenius or definition of acids, bases, and salts to explore the properties and use of the important chemical compounds whose water solutions contain ions. The important properties of an acid, a salt, and a base are shown; also shown are the ways in which they are prepared commercially and in the chemistry laboratory. (Color - 21 min., B/W) ACID-BASE INDICATORS, #4130 Interprets the experimental behavior of acid-base indicators using proton-donor/acceptor theory. Experiments and animated demonstration show effects of changing acidity on indicators. Mixing the indicators demonstrates proton competition among bases, with each changing color at different total acidity. (Color - 19 min.)

ACIDS, BASES, SALTS

CONTENT	LEARNING EXPERIENCES
E. Mechanism of Ionization 1. Structural formulas of acids a. Mineral b. Organic 2. Formation of H_3O^+ ion 3. Amphiprotic substances $\text{Al}(\text{OH})_3$ 4. Percentage of ionization a. Weak acids b. Weak bases c. Strong acids d. Strong bases	E-1 Demonstrate ball and stick models of structures of inorganic and organic acids. E-2 Discuss mechanism of formation of H_3O^+ ion using ionic equations. E-3 Discuss amphoterism. Discuss $\text{Al}(\text{OH})_3$ as acid and base. Discuss OH^- group in both acids and bases. E-4 Show weak acids and bases by means of 0.1M solutions and conductivity apparatus or with Gramercy Universal Indicator. E-5 Solubilities of hydroxides- Alyea 4-20
TEXT REFERENCES	AUDIO-VISUAL AIDS
Baker, Bradbury, et al., pp. 240-243 252-253 Metcalfe, Williams, Castka, pp. 218-222 228-229 Smoot, Price, Barrett, pp. 371-373	<u>FOM Filmstrip</u> ACID-BASE THEORIES, #636 Arrhenius and Bronsted concepts are developed by experiments and appropriate models. Conjugate acid-base table is explained. Indicator used related to pH. <u>Welch Scientific Company (Filmstrip)</u> ACIDS, BASES AND SALTS, #3992S, Group II

ACIDS, BASES, SALTS

CONTENT	LEARNING EXPERIENCES
F. pH 1. Definition 2. pH scale 3. Determination by indicators 4. Calculation from H^+ and OH^- concentrations	F-1 Demonstrate concept of pH values by measuring pH of common substances, such as H_2O, vinegar, milk of magnesia, 0.1M NaOH with pH paper, universal indicator and/or pH meter. F-2 Student experiment on determination of hydrogen ion concentration of a variety of acid and basic solutions using indicators. F-3 Homework assignment on pH calculation.
TEXT REFERENCES	AUDIO-VISUAL AIDS
Baker, Bradbury, et al., p. 282 Metcalf, Williams, Castka, pp. 234-238 Smoot, Price, Barrett, pp. 382-385	<u>McGraw-Hill</u> HYDROGEN ION CONCENTRATION, Set #6, #405490 <u>Chemistry Lab Series</u> INDICATORS AND PH The effect of hydrochloric acid solutions on indicators is shown, and the pH of a solution of acetic acid is determined. (Color - 28 min., B/W) TITRATING WITH PHENOLPHTHALEIN Illustrated and defines the range of phenolphthalein as an indicator. Demonstrates a correct technique of titrating and reaching an end point using phenolphthalein. 8mm -- Color -- #R80458 Super 8mm -- Color -- #S80458

ACIDS, BASES, SALTS

CONTENT	LEARNING EXPERIENCES
G. Salts 1. Definition 2. Formation a. By acid-base neutralization b. By direct combination of elements c. By acids and active metals d. By acids and metallic oxides e. By bases and nonmetallic oxides f. By metallic oxides and non-metallic oxides	G-2 Demonstrate salt formations by recording the salts formed in the reaction of $\text{HCl} + \text{NaOH}$, $\text{Zn} + \text{HCl}$, $\text{MgO} + \text{HCl}$. Precipitate CaCO_3 by blowing breath (CO_2) through limewater. Dissolve CaO and P_2O_5 in H_2O; evaporate.
TEXT REFERENCES	AUDIO-VISUAL AIDS
Baker, Bradbury, et al., pp. 253-254 Metcalf, Williams, Castka, pp. 244-246 Smoot, Price, Barrett, pp. 370-371	

ACIDS, BASES, SALTS

CONTENT	LEARNING EXPERIENCES
G (Continued) 3. Hydrolysis a. Definition b. Mechanism of hydrolysis i. Salt of strong acid-strong base ii. Salt of strong acid-weak base iii. Salt of weak acid-strong base iv. Salt of weak acid-weak base	G-3(i) Compare pH paper tests on solutions of Na_2CO_3 and NaHCO_3. Expand to use of NaHCO_3, but not Na_2CO_3 as an antacid. Use universal indicator on 0.1N solutions of $\text{Zn}(\text{NO}_3)_2$ - NaNO_3 - NH_4NO_3. G-3(ii) Construct student experiments to determine the effect of salt solutions on litmus or other indicators. G-3(iii) Student experiments to determine the degree of hydrolysis of 0.1M Na_2CO_3 using pH test paper.
TEXT REFERENCES	AUDIO-VISUAL AIDS
Baker, Bradbury, et al., pp. 315-319 Metcalf, Williams, Castka, pp. 367-368 Smoot, Price, Barrett, pp. 387-389	<u>McGraw-Hill</u> HYDROLYSIS OF SALTS, Set #5, #019100

COLLOIDS, COLLOIDAL SYSTEMS

THE APPROACH

Some modern texts in general chemistry do not have an extensive section on colloids. However, all make some mention of macromolecules (polymers) and surface phenomena (catalysis, adsorption). Because of the important role which colloids play in living systems, one or two periods may be spent introducing the main features of colloid chemistry.

Stress should be placed on the necessity for two phases - a colloidal is not a homogeneous substance but consists of two phases - the dispersed phase and the dispersion medium. Since all gases are miscible in each other, one of the nine possible combinations is eliminated (solid-solid, liquid-solid, gas-solid, etc.).

The "colloidal state" has often been referred to as the fourth state of matter, indicating the importance of particle size and the dual phase requirement. All three states of matter - solid, liquid, gas - may interact to appear in the colloidal state.

The study of colloids will normally follow the discussion of units on states of matter, solutions, and electrolytes. A brief review may also be included when organic compounds are studied.

The Journal of Chemical Education published a number of papers on the Teaching of Colloid and Surface Chemistry in 1962, J. Ch. Ed. 39-167 (1962).

OBJECTIVES

1. The student should be able to state the differences between a colloidal suspension and a solution, with respect to the following characteristics:
 - a. particle size
 - b. phases present
 - c. permanence
 - d. light scattering (Tyndall effect)
 - e. dialysis
2. The student should be able to list six to eight types of colloidal suspensions.
3. The student should be able to define the process of adsorption as it applies to colloidal suspensions.
4. The student should be able to define the Tyndall effect as it applies to colloidal suspensions.

CC716-2

5. The student should be able to define the process of electrophoresis as it applies to a colloidal suspension.
6. The student should be able to define each of the following methods of preparing colloidal suspensions:
 - a. homogenization
 - b. peptization
 - c. emulsification

COLLOIDS, COLLOIDAL SYSTEMS

CONTENT	LEARNING EXPERIENCES
A. Definition 1. Dispersed phase 2. Dispersion medium B. Size of Particles 1. Molecular (less 10^{-7} cm) (invisible) 2. Colloidal (ultramicroscope) 3. Precipitate (greater than 10^{-5} cm) (microscope or naked eye) C. Types of Systems 1. Liquid in gas (fog, mist) 2. Solid in gas (smoke) 3. Gas in liquid (foam) 4. Liquid in liquid (emulsion milk) 5. Solid in liquid (sol or gel) 6. Gas in solid (solid foam-pumice) 7. Liquid in solid (opal) 8. Solid in solid (red glass)	A-B Experiments on types, size, and properties in Alyea C-1 Contrails, fog in valley C-2 Cigars, pipe smoke, fume NH_3 near HCl C-3 Soap suds C-4 Two bottles of H_2O and kerosene. Add little soap in bottle and shake both. Compare. C-5 Dissolve gelatine in hot H_2O and cool. C-6 Piece of foam rubber C-7 Piece of milky quartz
TEXT REFERENCES	AUDIO-VISUAL AIDS
Baker, Bradbury, et al., pp. 637-690 692-293 Metcalf Williams, Castka, pp.251-255 Smoot, Price, Barrett, pp. 526-530	<u>FOM Filmstrip</u> COLLOIDS. #665 The size relationship of dispersed particles in the colloidal state is discussed, as well as the significant properties of colloidal suspensions such as settling out and performance, filterability, passage, light scattering, Brownian movement, adsorption, electrical charge, and colligative properties. Illustrates relationships of colloidal state to industry, technology and biology. <u>Welsh Scientific Company</u> Chemistry for Today: Group I Filmstrip SOLUTIONS, SUSPENSIONS AND COLLOIDS, #3992R

COLLOIDS, COLLOIDAL SYSTEMS

CONTENT	LEARNING EXPERIENCES
D. Preparation of Systems 1. Condensation of small particles a. Controlled precipitation b. Ppt. in presence of protecting colloid 2. Dispersion of larger particles a. Peptization b. Use of colloid mill or homogenizer c. Arcing metal in liquid E. Properties 1. Motion of particles (Brownism) 2. Light scattering (Tyndall) 3. Color related to size 4. Electrical charge (Electrophoresis) 5. Stability (flocculation) 6. Surface energy (adsorption)	D-1 Ppt. Ag using ammoniacal AgNO_3 solution and a trace of formaldehyde. a. Add FeCl_3 solution drop by drop to boiling H_2O. b. Add H_2O to baking powder + a bit of albumin. D-2 b. Colloid mills in scientific catalogs c. Strike an arc under H_2O using Ag electrodes. E-1 Place pollen in H_2O under microscope. E-2 Shoot beam of light through a solution of hypo, then slowly add HCl to solution. E-3 Vary concentrations in Ex. D-1a. E-4 U tube with agar containing small amount of $\text{K}_4\text{Fe}(\text{CN})_6$ and 5 drops of phenolphthalein. Put FeCl_3 solution on anode side and pass current through.
TEXT REFERENCES	AUDIO-VISUAL AIDS
Baker, Bradbury, et al., pp. 699-705 Metcalf, Williams, Castka, pp. 184 254-259 Smoot, Price, Barrett, pp. 529, 532-534	<u>Coronet Films</u> THE COLLOIDAL STATE The products of colloidal matter are of major significance in a multitude of chemical processes. This film introduces us to the colloidal state through a series of carefully controlled laboratory experiments and demonstrations. It defines colloids, and their several kinds, shows how they differ from solutions and suspensions, how they may be prepared and destroyed, and points out the many uses of colloids in the chemistry of everyday life. (Color - 16 min. B/W)

OXIDATION-REDUCTION

THE APPROACH

In an early section of the usual chemistry program, oxidation-reduction is treated in a qualitative way in order to explain burning and simple reactions of hydrogen and oxygen. Thus, redox is demonstrated as merely an addition or subtraction of oxygen. In the beginning of a chemistry course, when students are somewhat timid and slightly confused, such a treatment is not unwise. It may be appropriate to postpone this more extensive discussion of redox theory until the unit on ionization.

The unit may be introduced with a short demonstration in which warmed sodium metal is lowered into a gas bottle of chlorine gas. When the sodium bursts into flame, the students are asked, "What kind of reaction is this?" Most will answer burning or combustion. "If combustion is rapid oxidation with heat and light emitted, how can this reaction be combustion without oxygen?" It can be, of course, if combustion or oxidation can be defined in terms of some concept other than oxygen transfer.

Memorization of rules for determining the oxidation number of elements in compounds is suggested, followed by practice and drill in finding oxidation numbers. Once the students have become proficient in assigning oxidation numbers, then simple equations may be assigned, followed by more difficult equations. If the class consists of a fast-learning or science-prone group, a few equations of significant difficulty may be assigned. Later, the instructor may wish to assign ionic equations to be balanced by the ion-electron, half-reaction method.

SOME TYPICAL UNBALANCED REDOX REACTIONS

Easy types

1. $\text{Mg} + \text{O}_2 \rightarrow \text{MgO}$
2. $\text{KClO}_3 \rightarrow \text{KCl} + \text{O}_2$
3. $\text{Cl}_2 + \text{NaBr} \rightarrow \text{NaCl} + \text{Br}_2$
4. $\text{Al} + \text{H}_3\text{PO}_4 \rightarrow \text{AlPO}_4 + \text{H}_2$

Moderate Difficulty

1. $\text{CdS} + \text{I}_2 + \text{HCl} \rightarrow \text{CdCl}_2 + \text{HI} + \text{S}$
2. $\text{MnO}_2 + \text{HCl} \rightarrow \text{MnCl}_2 + \text{Cl}_2 + \text{H}_2\text{O}$
3. $\text{HNO}_3 + \text{H}_2\text{S} \rightarrow \text{NO} + \text{S} + \text{H}_2\text{O}$
4. $\text{Cu} + \text{HNO}_3 \rightarrow \text{Cu}(\text{NO}_3)_2 + \text{NO} + \text{H}_2\text{O}$

Very Difficult

1. $\text{Cl}_2 + \text{H}_2\text{O} \rightarrow \text{HCl} + \text{HClO}$ (Auto-oxidation)
2. $\text{KMnO}_4 + \text{H}_2\text{O}_2 + \text{H}_2\text{SO}_4 \rightarrow \text{KHSO}_4 + \text{H}_2\text{O} + \text{MnSO}_4 + \text{O}_2$
3. $\text{KOH} + \text{Cl}_2 + \text{CrI}_3 \rightarrow \text{K}_2\text{CrO}_4 + \text{KIO}_4 + \text{KCl} + \text{H}_2\text{O}$
4. $\text{As}_2\text{S}_5 + \text{HNO}_3 \rightarrow \text{H}_3\text{AsO}_4 + \text{H}_2\text{O} + \text{NO}_2$

If time permits, normality of reagents in redox reactions may be discussed. The experiment "Oxidation-reduction Titration", in which a ferrous salt (such as FeSO_4) is titrated against a standard KMnO_4 solution, may be an effective demonstration.

OBJECTIVES

1. The student should be able to define the processes of oxidation and reduction in two ways.
2. The student should be able to give a specific example of each definition using chemical equations.
3. The student should be able to identify, in a given equation, the following:
 - a. oxidizing agent
 - b. reducing agent
 - c. oxidized element
 - d. reduced element
4. The student should be able to determine the oxidation number of an element in a compound, given the formula of the compound.
5. The student should be able to identify the oxidation numbers of eight out of ten elements from the electronic configuration of the element.
6. The student should be able to balance eight out of ten redox equations using the electron transfer method.
7. The student should be able to determine the gram-equivalent weight of an element which is involved in a redox reaction.

OXIDATION-REDUCTION

CONTENT	LEARNING EXPERIENCES
A. <u>Definitions</u> 1. Narrow concept of gain-loss of oxygen 2. General concept of gain-loss of electrons	A-1 Compare reactions of burning carbon in oxygen and burning hydrogen in chlorine to show relationship of old and new concepts. A-2 Stress simultaneous occurrence of the two processes.
TEXT REFERENCES	AUDIO-VISUAL AIDS
Baker, Bradbury, <u>et al.</u>, pp. 72, 351 Metcalf, Williams, Castka, pp. 377-378 Smoot, Price, Barrett, pp. 400-401	NE PA A-V LIBRARY <u>McGraw-Hill</u> OXIDATION-REDUCTION, # A-823 (9 1/2 min. - Color) Examples of oxidation and reduction demonstrate the complementary relationship and the differences between the two. A basic and broadened explanation of this type of reaction in terms of electron transfer is provided.

OXIDATION-REDUCTION

CONTENT	LEARNING EXPERIENCES
B. <u>Oxidizing-Reducing Agents</u> 1. Definition 2. Relative strengths 3. Auto-oxidation	B-1 Demonstrate with an oxidation-reduction reaction such as the oxidation of glucose by the oxygen in H_2O (KOH solution) which is made visible by methylene blue. (Modern Chemistry Teachers Manual, p. 190, or Alyea 8-15, 8-17). Very dramatic. B-2 Student experiment on oxidation-reduction such as oxidation of copper by nitric acid, evaporation to dryness, conversion of the copper nitrate into CuO, and subsequent reduction of the oxide to copper by carbon. (Scientific Experiments in Chemistry, MCA-Weaver).
TEXT REFERENCES	AUDIO-VISUAL AIDS
Baker, Bradbury, <u>et al.</u>, pp. 105, 306 Metcalf, Williams, Castka, p. 378 Smoot, Price, Barrett, pp. 403-404	

OXIDATION-REDUCTION

CONTENT	LEARNING EXPERIENCES
C. <u>Oxidation Number</u> 1. Oxidation number and relationship to valency theory 2. Relationship to electronic configuration 3. Assignment of chemical formulas D. <u>Balancing redox equations</u> 1. Assign oxidation numbers 2. Show electron transfer in half-reactions 3. Balance electron transfer 4. Balance skeleton equation 5. Balance entire equation	C Drill on assignment of oxidation numbers in compounds. D-1 Develop steps involved in balancing redox equations. Illustrate the procedure with a few examples in order of increasing difficulty, such as: a) $\text{H}_2\text{S} + \text{O}_2$ b) $\text{MnO}_2 + \text{HCl}$ c) $\text{HCl} + \text{KMnO}_4$ D-2 Homework assignment on balancing equations by redox method.
TEXT REFERENCES	AUDIO-VISUAL AIDS
Baker, Bradbury, <u>et al</u>, pp. 105-106, 351-356. Metcalfe, Williams, Castka, pp. 378-385 Smoot, Price, Barrett, pp. 404-407, 407-415	<u>McGraw-Hill Filmstrip</u> BALANCING EQUATIONS BY ELECTRON TRANSFER, Set No. 4, # 019200 (Color) <u>FOM Filmstrip</u> OXIDATION-REDUCTION, # 642 Rules for the determination of oxidation state and oxidation number are explained and applied. Various methods for balancing redox reactions include electron-shift, ion-electron, and half-reactions.

CHEMICAL KINETICS

THE APPROACH

In the beginning it is important to communicate to the student the meaning of the term chemical kinetics. Chemical kinetics is concerned with the sequence of steps by which chemical reactions occur and the rate at which each step occurs. At an early stage in the development of this topic, heats of reaction should be clearly defined and the significance of exothermic and endothermic reactions explained. This includes writing thermochemical equations. The symbol ΔH_{rx} (heat content or enthalpy of products - heat content or enthalpy of reactants) is used for denoting the heat of reaction. Clarify the arbitrary sign convention of ΔH for endothermic and exothermic reactions.

A fundamental quantity of reactant heats, the heat of formation, should now be introduced. It may be pointed out that most compounds have negative heats of formation and only a few, such as hydrogen iodide and carbon disulfide, have positive heats of formation. The reason for this may be explored at this point. A good example to use is the formation of water molecules from the nonpolar diatomic molecules of hydrogen and oxygen. Before concluding the discussion of heats of formation, show how this property of a compound is related to its stability.

Another category of reaction heats i. e., heats of combustion, may be introduced at this point. The term, heat of combustion, should be clearly defined and a number of simple examples illustrated. Show how the heat of combustion of a substance can be a useful tool in determining the heat of formation of a substance which cannot be formed in a composition reaction directly from its constituent elements. For example, the complete combustion products of many organic compounds are CO_2 and H_2O . Since the heats of formation of these two substances are known, the heats of formation of the organic compounds can be calculated from the following equation

$$\begin{array}{lcl} \text{Heat of formation} & = & \text{Sum of heats of formation} \\ \text{of compound X} & & \text{of products of combustion} \\ & & \text{of compound X} \end{array} \quad \begin{array}{l} \text{Heat of combustion} \\ \text{of compound X} \end{array}$$

It should be emphasized that the change in heat content of a reaction system is related to the change in the number and strength of bonds as the system transforms from reactants to products. The reaction for the formation of water gas is a good example to use in order to illustrate this relationship.

At this point the driving force of reactions should be introduced and explained. A quantitative approach to this topic should not be attempted at this point. However, it is within the realm of an ordinary high school class to examine qualitatively the "driving force" of chemical reactions. Factors such as the tendency for processes to occur which lead to the lowest possible state of

energy and a tendency for processes to occur which leads to the highest possible state of disorder or increased entropy can be examined in some detail. In cases where the energy change and the entropy change oppose each other, it should be mentioned that the system will go in the direction of the larger change. The net driving force is called the free energy change of the system. At constant atmospheric pressure and constant temperature the free energy change is denoted by $\Delta G = \Delta H - T \Delta S$, where ΔG is the change in free energy of the system, ΔH is the change in heat content or enthalpy, T is the temperature in °K, and ΔS is the change in entropy. In this case ΔS is multiplied by T to give the term the same unit dimensions as ΔH . A chemical reaction will proceed if there is a decrease in free energy (negative ΔG) between reactants and products. This implies that free energy is released to do useful work.

At this point activation energy might be explained. A simple plot utilizing energy content of reactants, energy of reaction and energy content of products may be introduced. Activation energy may be shown on plot. A plot may then be made with ΔG and free energy changes illustrated including ΔG_{act} .

At this point it becomes necessary to consider specific factors which influence the rate of chemical reaction. These factors include:

- (1) Nature of the reactants
- (2) Amount of surface area
- (3) Concentration of reactants
- (4) Temperature
- (5) Action of catalysts
- (6) Purity of reactants

Each of these factors should be investigated independently and it should be shown that a change which affects the reaction rate is a change in conditions which affects either the frequency of collisions or the average energy of collisions.

Before concluding this unit, the Law of Mass Action should be explained. It should be emphasized that the Law of Mass Action as stated in most texts is directly applicable to an ideal homogeneous reaction which occurs in one step. For more complex reactions the mass-action principle cannot be applied unless modified.

OBJECTIVES

1. The student should be able to give an acceptable definition for the term kinetics.
2. The student should be able to solve 7 out of 10 problems determining ΔH_{rx} for a given reaction.

3. The student should be able to interpret either a positive ΔH or a negative ΔH as being either an exothermic or endothermic reaction enthalpy.
4. The student should be able to list the order of stabilities of 5 compounds from heat of formation table 3.
5. The student should be able to determine the heat of formation of 5 compounds using the ΔH_{rx} equation.
6. The student should be able to solve 4 out of 5 problems determining the heat of combustion of a compound from heats of formation.
7. The student should be able to list 6 factors that influence the rate of a chemical reaction.
8. The student should be able to state the difference in activation energy as it applies to an exothermic reaction and as it applies to an endothermic reaction.
9. The student should be able to define each of the following types of equilibrium:
a. physical b. chemical c. static d. dynamic
10. The student should be able to draw a graph of potential energy vs. time for a given chemical reaction showing how activation energy is illustrated in the graph.
11. In the graph mentioned in Objective No. 10, the student should be able to label the position of the activated complex.
12. The student should be able to plot a diagram of potential energy vs. time for an exothermic reaction.
13. The student should be able to plot a diagram of potential energy vs. time for an endothermic reaction.
14. The student should be able to solve 4 out of 5 problems using the K_{eq} equation.
15. The student should be able to state the Law of Mass Action equation.
16. The student should be able to write the rate expression for 4 out of 5 simple chemical reactions.
17. The student should be able to define entropy.
18. The student should be able to define free energy.
19. The student should be able to state the relationship between free energy change and a spontaneous reaction.

CHEMICAL KINETICS

CONTENT	LEARNING EXPERIENCES
A. Introduction 1. Definition of kinetics 2. Energy changes in chemical reactions a. Heats of reaction i. Designation of sign of ΔH_{rx} ii. Relationship to bond energies b. Heats of formation c. Heats of combustion	Student experiment to determine the heat liberated by a neutralization reaction. Stress sign convention with values for heat of formation.
TEXT REFERENCES	AUDIO-VISUAL AIDS
Baker, Bradbury, et al., pp. 78-79; 294-295 Metcalfe, Williams, Castka, pp. 332-338 Smoot, Price, Barrett, pp. 94-95; 96-97	<u>Northeastern Pennsylvania Audio-Visual Center</u> <u>EXPLAINING MATTER: CHEMICAL CHANGE,</u> <u>#A-818, (11 min. - Color)</u>

CHEMICAL KINETICS

CONTENT	LEARNING EXPERIENCES
B. Factors Influencing Reaction Rate 1. Nature of reactants 2. Exposed surface area 3. Concentration of reactants 4. Purity of reactants 5. Temperature 6. Catalysts	Review factors affecting rate of solution; correlate with reaction rate. B-1 Demonstrate effect of nature of reactants by comparing rate of precipitation of AgNO_3 + HCl with that of acidified Na_2HPO_4 + $(\text{NH}_4)_2\text{MoO}_4$. B-2 Demonstrate effect of surface area by pouring dilute H_2SO_4, HNO_3 and HCl upon marble chips. H_2SO_4 reaction stops as insoluble CaSO_4 covers the chips. B-3 Student experiment to determine effect of concentration on reaction rate using constant weight of Mg ribbon and varying concentration of HCl. B-5 Student experiment to determine the effect of temperature on reaction rate by timing reaction of KIO_3 with acidified NaHSO_3 - starch solution at various temperatures.
TEXT REFERENCES	AUDIO-VISUAL AIDS
Baker, Bradbury, et al., pp. 302-307 Metcalf, Williams, Castka, pp. 346-350 Smoot, Price, Barrett, pp. 340-348	McGraw-Hill Filmstrip, Set #5 <u>RATE OF REACTION</u> <u>FOM Filmstrip</u> <u>CATALYSTS AT WORK, #609</u> Two main types of catalysts are explained in terms of numerous experiments in which are reproduced in the classroom. Many catalytic industrial processes are illustrated. <u>Northeastern Pennsylvania Audio-Visual Center</u> <u>CATALYSIS, #1008 (17 min. - Color)</u> <u>Chem Lab Series</u> <u>RATE OF REACTION</u> Several factors which affect reaction rates are discussed and illustrated. The reaction between crystals of potassium permanganate and glycerine show the effect of temperature. The reaction of oxalic acid solutions and permanganate solutions at various temperatures shows the effect of temperature. Other demonstrations illustrate additional factors which influence reaction rate. (28 min. - \$125--B/W - \$250--Color)

CHEMICAL KINETICS

CONTENT	LEARNING EXPERIENCES
C. Collision Theory 1. "Reshuffling of chemical bonds" 2. Activation energy 3. Activated complex 4. Rate of reaction = number of effective collisions / time.	C-2 Indicate or illustrate how activation energy may be of different forms, i.e., heat in a match and light is activation source in photography. Demonstration - place hydrogen and oxygen in a container. No evidence of a reaction in a sealed container no matter how long they stand. Then ignite the mixture with a spark and observe how the gases react chemically. Indicate activation energy in everyday life. For example, eggs are heated so that protein will react to form a more palatable food.
TEXT REFERENCES	AUDIO-VISUAL AIDS
Baker, Bradbury, et al., pp. 296-297 Metcalf, Williams, Castka, pp. 342-346 Smoot, Price, Barrett, pp. 341-344	<u>MLA</u> INTRODUCTION TO REACTION KINETICS . #4121 Illustrates the mechanisms of some simple chemical reactions. Explains the effect of temperature, activation energy, geometry of collision, and catalysis upon the rate of reaction. The reactions between hydrogen and iodine are used. The speed of the activation has been slowed down by a factor of 10^{14}. Potential energy curves clarify the relationship between the energy required for a reaction to occur and the relative position of the reaction particles before, during, and after the collision. (13 min. - Color) ENERGY AND REACTION The relationship between energy, in its various manifestations, and chemical reactions is shown. This is then related to such fundamental concepts as the making and breaking of chemical bonds, activation energy, and rate of reaction. (14 1/2 min. - Color #612023- B/W #612011)

CHEMICAL KINETICS

CONTENT	LEARNING EXPERIENCES																
	C-3 Illustrate using reaction between H_2 and O_2. Show how kinetic energy of motion is transformed into the potential energy of reaction. This example can be used to define an activated complex.																
TEXT REFERENCES	AUDIO-VISUAL AIDS																
	<u>FOM Filmstrips</u> CHEMICAL KINETICS, #6500 Various factors involved in rates of reaction and their measurement explained. Details of related collision theory, activated complex and a variety of specific reactions utilized in exposition of reaction mechanisms. <u>PSP Transparencies, Set TC-1</u> CHEMICAL KINETICS Reaction rates explained in terms of activation energy, reaction mechanisms and reaction pathways. Reaction rate factors such as energy, orientation and concentration of reactants are illustrated by means of graphs and diagrams. Shows the effect of catalysts in generating new reaction pathways. <table data-bbox="799 2069 1808 2499"> <thead> <tr> <th>Transparency Title</th><th>Overlays</th></tr> </thead> <tbody> <tr> <td>Collision Theory</td><td>2</td></tr> <tr> <td>Collision Frequency and Concentration</td><td>2</td></tr> <tr> <td>Temperature, Kinetic Energy, Activation Energy</td><td>2</td></tr> <tr> <td>Reaction Pathways: Uncatalyzed and Catlyzed</td><td>2</td></tr> <tr> <td>A Reaction Mechanism (Formation of HCl)</td><td>2</td></tr> <tr> <td>Reaching Equilibrium</td><td>2</td></tr> <tr> <td>Reaction Rate Factors</td><td>2</td></tr> </tbody> </table>	Transparency Title	Overlays	Collision Theory	2	Collision Frequency and Concentration	2	Temperature, Kinetic Energy, Activation Energy	2	Reaction Pathways: Uncatalyzed and Catlyzed	2	A Reaction Mechanism (Formation of HCl)	2	Reaching Equilibrium	2	Reaction Rate Factors	2
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CHEMICAL KINETICS

CONTENT	LEARNING EXPERIENCES
C. (Continued) 5. Potential energy versus reaction diagrams a. For exothermic reaction b. For endothermic reaction c. Activated complex and activation energy d. Effect of temperature e. Effect of catalysts	5(a) Plot an actual diagram using the reaction between H_2 and O_2 to show an exothermic reaction. 5(b) Plot an actual diagram using the decomposition of H_2O to show an endothermic reaction. [Note: This diagram is the opposite of 5(a)] 5(c) Illustrate activated complex and activation energy. 5(d) Emphasize that an increase in temperature will increase the rate of a reaction because more molecules become more energetic and more activated complexes are formed. In other words the increased number of collisions are effective
TEXT REFERENCES	AUDIO-VISUAL AIDS
Baker, Bradbury, et al., pp. 297;305-307 Metcalfe, Williams, Castka, pp. 342-343 Smoot, Price, Barrett, pp. 341-345	<u>MLA</u> CATALYSIS, # 4127 Collaborator: Professor Richard E. Powell, University of California, Berkeley, California. Produced in cooperation with the Manufacturing Chemists' Association. Emphasizes that catalysts are typical chemical reactants, unique only in that they are regenerated during reaction. Animation shows what takes place on the molecular level in a catalyzed reaction. Potential energy curves give the relationship between uncatalyzed and catalyzed reactions. (17 min Color - \$ 135)

CHEMICAL KINETICS

CONTENT	LEARNING EXPERIENCES
D. Equilibrium 1. Physical 2. Chemical 3. Static 4. Dynamic	D-4 Use evaporation in a closed container as an example of dynamic equilibrium. Indicate that the point when as many particles are returning to the surface as are leaving is dynamic equilibrium.
TEXT REFERENCES	AUDIO-VISUAL AIDS
Baker, Bradbury, et al., pp. 308-309 Metcalf, Williams, Castka, pp. 355-357 Smoot, Price, Barrett, pp. 227-228	<u>McGraw-Hill Filmstrip, Set #5</u> EQUILIBRIUM, #019210 <u>Northeastern Pennsylvania Audio-Visual Center</u> EQUILIBRIUM, #A-802, (24 min. - Color)

CHEMICAL KINETICS

CONTENT	LEARNING EXPERIENCES
E. Law of Mass Action 1. $aA + bB \rightleftharpoons cC + dD$ 2. $\text{Speed}_1 = k_1 (A)^a \times (B)^b$ 3. $\text{Speed}_2 = k_2 (C)^c \times (D)^d$ 4. $\text{Speed}_1 = \text{Speed}_2$ at equilibrium 5. $K_{eq} = \frac{k_1}{k_2} = \frac{(C)^c \times (D)^d}{(A)^a \times (B)^b}$	E(2-3) Demonstrate by reacting SO_2 and O_2 to form SO_3. Compare rate of reaction when the concentration of O_2 is doubled. Compare this result with theoretical result obtained when this problem is solved using the Law of Mass Action equation. E(4-5) Show the technique for solving problems involving equilibrium constant, i.e. $K_{eq} = \frac{K_f}{K_b}$ where f refers to forward reaction and b refers to reverse reaction.
TEXT REFERENCES	AUDIO-VISUAL AIDS
Baker, Bradbury, et al., pp. 320-322 Metcalf, Williams, Castka, pp. 350-351 Smoot, Price, Barrett, pp. 348-351	McGraw-Hill Filmstrip, Set #6 EQUILIBRIUM CONSTANT -K_{eq}, #405490

CHEMICAL KINETICS

CONTENT	LEARNING EXPERIENCES
F. Probability of a Chemical Reaction 1. Relationship of enthalpy change, ΔH 2. Relationship of change in degree of disorder a. Definition of entropy b. ΔS	F-1 Assign a number of chemical reactions and have the class predict the relative tendencies of these reactions to occur by determining the ΔH values for each reaction. F-2 A simple demonstration involving the melting of ice can be used to illustrate a positive ΔS as the system goes from greater order to less order. The solution of a gas can be used to illustrate the opposite process, i.e., a negative ΔS.
TEXT REFERENCES	AUDIO-VISUAL AIDS
Baker, Bradbury, et al. _____ Metcalfe, Williams, Castka, pp. 38; 339-341 Smoot, Price, Barrett, pp. 338-339	

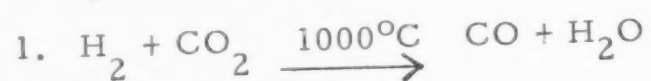
CHEMICAL KINETICS

CONTENT	LEARNING EXPERIENCES
F. (Continued) 3. Free energy a. Definition b. $\Delta G = \Delta H - T \Delta S$	3(b) Indicate that a negative free energy change for a reaction indicates it will occur spontaneously and vice versa.
TEXT REFERENCES	AUDIO-VISUAL AIDS
Baker, Bradbury, et al. ——— Metcalfe, Williams, Castka, p. 341 Smoot, Price, Barrett, pp. 338-339	121

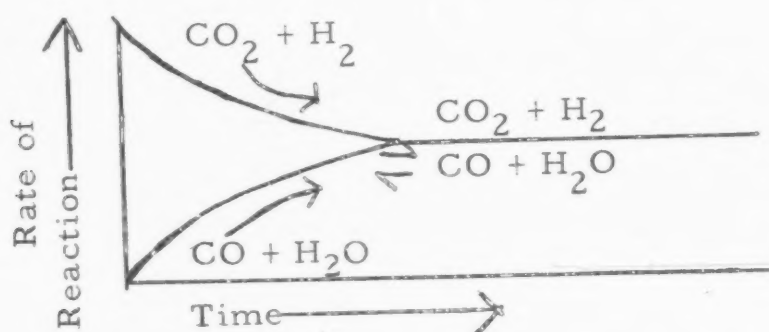
EQUILIBRIUM

THE APPROACH

The unit on chemical equilibrium may be initiated by a discussion of the factors which influence the composition of the equilibrium system. The concept of reversible reactions and their direct relation to chemical equilibrium should be introduced early. To illustrate the relationship which exists, the following reactions can be given as illustrations:



The above equations and their relation to reversibility and equilibrium can then be explained in terms of the following graphic illustration:



To show that the equilibrium state is dynamic rather than static, it should be explained that when ordinary water is mixed with heavy water, D_2O , and an equilibrium has been established, some D_2 and DH are found in the system.

It should be emphasized that in an equilibrium system, the rate of the forward and reverse reactions is equal, but the quantity of products and reactants is not.

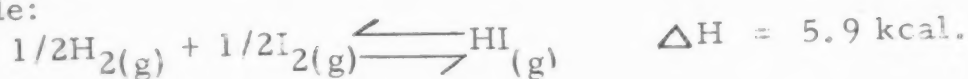
A good approach is to separate the discussion into one dealing with the qualitative aspects of equilibrium and one dealing with the quantitative aspects of equilibrium.

To initiate a qualitative discussion of this topic, LeChatelier's Principle should now be defined and thoroughly explained. The following three factors which may alter the composition of a system at equilibrium should now be discussed individually:

1. Concentration
2. Pressure
3. Temperature

at equilibrium, the reaction which takes place with the absorption of heat will be promoted.

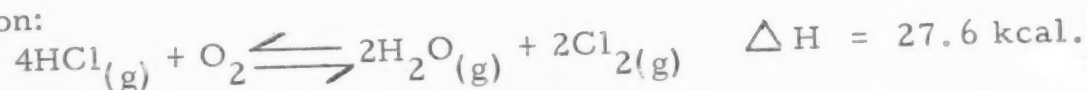
Example:



An increase in temperature will produce more HI because of the positive ΔH (ΔH is positive when heat is absorbed and negative when heat is evolved.) It may be necessary at this point to review heats of reactions.

Ask the class to answer the following question and explain the answer: What happens to the yield of chlorine in the following reaction if the temperature is increased?

Equation:



At this juncture the student should be ready to discuss the quantitative treatment of chemical equilibrium. To initiate the discussion the equilibrium constant should be defined and its significance should be explained. Definition: The equilibrium constant is equal to the ratio of the products of the concentrations of the substances produced to the concentrations of the reacting substances with each concentration raised to a power equal to its numerical coefficient in the balanced equation.

Example:



$$\text{Forward reaction rate (1)} = K_1 \cdot C_{\text{H}_2} \cdot C_{\text{I}_2}$$

$$\text{Reverse reaction rate (2)} = K_2 \cdot C_{\text{HI}}^2$$

K_1 is a proportionality constant which is characteristic of the reaction and is defined as the specific reaction rate constant. It is not affected by concentration but may change with temperature. At equilibrium, Rate (1) = Rate (2) so that $K_1 \cdot C_{\text{H}_2} \cdot C_{\text{I}_2} = K_2 \cdot C_{\text{HI}}^2$ (concentrations are usually expressed in moles per liter.)

As a result of the previous rate expression, we can now state the following equation:

$$\frac{(\text{HI})^2}{(\text{H}_2)(\text{I}_2)} = \frac{K_1}{K_2} = K_e \quad (\text{Equilibrium Constant})$$

The above equation is referred to as the Law of Chemical Equilibrium. It should now be pointed out that all reversible reactions eventually attain a concentration quotient, Q_e , which is equal to their respective K_e , irrespective of the composition of the starting mixture. If one begins with pure HI, the ratio of the concentration of H_2 to that of I_2 produced will always be unity. However, irrespective of the ratio of H_2 to I_2 initially, at equilibrium, a concentration quotient is produced which is equal to K_e . Even if the concentration ratio, Q_e , is destroyed by adding H_2 to the system (this would make the concentration ratio, Q , less than Q_e), H_2 will react with I_2 to form HI until Q again becomes equal to Q_e and, hence, K_e .

It should be made clear that the magnitude of K_e is a measure of the extent to which a reaction occurs. A small value of K_e implies a small degree of reaction and a large value of K_e implies a considerable degree of reaction.

A specific reaction can now be written and the quantitative effect of various factors on the equilibrium determined.



$$\frac{(NH_3)^2}{(N_2)(H_2)^3} = K_e$$

If	Then
1. C_{H_2} is increased	C_{N_2} <u>decreases</u> , C_{NH_3} <u>increases</u>
2. C_{N_2} is increased	C_{H_2} <u>decreases</u> , C_{NH_3} <u>increases</u>
3. C_{NH_3} is increased	C_{N_2} <u>increases</u> , C_{H_2} <u>increases</u>
4. Pressure of system is increased	C_{NH_3} <u>increases</u> , C_{N_2} <u>decreases</u> , C_{H_2} <u>decreases</u>
5. C_{N_2} is decreased	C_{H_2} <u>increases</u> , C_{NH_3} <u>decreases</u>
6. C_{NH_3} is decreased	C_{N_2} <u>decreases</u> , C_{H_2} <u>decreases</u>
7. Volume of the system is increased	C_{NH_3} <u>decreases</u> , C_{N_2} <u>increases</u> , C_{H_2} <u>increases</u>
8. The system is heated	C_{NH_3} <u>decreases</u> , C_{N_2} <u>increases</u> , K_e <u>decreases</u> C_{H_2} <u>increases</u>

The answers portion in the right hand column can be left blank and serve as a good evaluative exercise for the student.

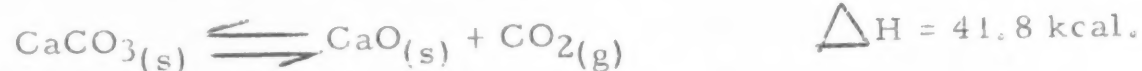
It should be emphasized that of all factors which effect equilibrium concentration, temperature is the only one which also affects the equilibrium constant. LeChatelier's Principle enables us to predict how K_e will vary with temperature.

For example, if heat is evolved in a reaction, a rise in temperature will decrease the extent of the reaction; Q_e decreases, hence K_e must become smaller. It follows that for an exothermic reaction K_e must decrease if the temperature increases.

*Exercise: Determine how K_e would vary for an endothermic reaction.

Before closing the unit on equilibrium, it should be shown that a solid in a reaction does not affect the equilibrium constant as long as some is present. It may therefore be ignored in writing the equilibrium constant expression. The reason is that the concentration of a solid is a constant at constant temperature, i.e. its density.

Example:



$$\frac{(\text{CaO})(\text{CO}_2)}{(\text{CaCO}_3)} = K_e$$

Increasing the concentration of CO_2 and hence its pressure, causes more $\text{CaCO}_{3(s)}$ to form. Removing CO_2 from the system causes $\text{CaCO}_{3(s)}$ to decompose. Experimentally, at a given temperature, K_e is found not to be affected by the addition or removal of $\text{CaO}_{(s)}$ or $\text{CaCO}_{3(s)}$. As long as some of each is present, K_e remains constant.

Since these solids do not affect K_e , their concentrations may be regarded as constants, so the equilibrium expression may be written:

$$\frac{k_1 \times (\text{CO}_2)}{k_2} = K \quad \text{or} \quad (\text{CO}_2) = \frac{K \times k_2}{k_1} = K_e$$

Sample problem: Write the equilibrium constant equation for $\text{C} + \text{O}_2 \rightleftharpoons \text{CO}_2$.

$$\text{Answer: } \frac{(\text{CO}_2)}{(\text{O}_2)} = K_e$$

The concentration of carbon does not enter into the equilibrium constant expression because it is a solid.

OBJECTIVES

1. The student should be able to define and illustrate a reversible reaction.
2. The student should be able to define chemical equilibrium.
3. The student should be able to give specific evidence that the equilibrium state is dynamic rather than static.
4. The student should be able to state 3 properties which may alter the composition of a system at equilibrium.
5. The student should be able to state LeChatelier's Principle.
6. Given a system at equilibrium, the student should be able to explain any changes which occur if the concentration of a component of the system is changed.
7. The student should be able to state 3 ways by which the concentration of components in an equilibrium system may be changed and affect the equilibrium.
8. Given a gaseous system at equilibrium, the student should be able to explain how the yield of one of the gases can be increased or decreased by a change in pressure on the system.
9. The student should be able to explain why the yield of a reaction will not be affected by pressure changes in certain gaseous equilibrium systems.
10. The student should be able to explain why pressure changes have little or no effect on liquid or solid systems.
11. The student should be able to explain the effect of a temperature increase on an equilibrium reaction with a positive Enthalpy change (ΔH).
12. The student should be able to explain the effect of a temperature increase on an equilibrium reaction with a negative Enthalpy change (ΔH).
13. Given 5 reactions at equilibrium, the student should be able to predict correctly how the yield of any one species will be affected by a change in either concentration of the components, pressure, or temperature.
14. The student should be able to define the equilibrium constant.
15. The student should be able to define the specific reaction rate constant.
16. The student should be able to state the significance of the specific reaction rate constant in terms of the equilibrium system.
17. The student should be able to write the mass action expression for a given equilibrium reactions.
18. The student should be able to state specifically what is implied by a value of K_e less than 1, and a value of K_e greater than 1.
19. The student should be able to state which factor affects the equilibrium constant.
20. The student should be able to tell why the concentration of solids does not affect the equilibrium constant.
21. The student should be able to solve 7 out of 10 problems of the following types given the equilibrium constant:
 - a. Determine equilibrium concentrations for all components in the system when one of the species is not present initially.
 - b. Determine equilibrium concentrations for all components in the system when all species are introduced into the system initially.

EQUILIBRIUM

CONTENT	LEARNING EXPERIENCES
A. Definition of Equilibrium 1. Physical 2. Chemical 3. Static 4. Dynamic	Demonstration on the dynamic nature of chemical equilibrium: <u>Chemistry</u>, Vol.41, #7, July-August, 1968, p. 48. A-4 Demonstrate dynamic equilibrium using a large beaker and a bell jar. <u>Chemistry-A Modern Course</u>, p. 228.
TEXT REFERENCES	AUDIO-VISUAL AIDS
Baker, Bradbury, et al., pp. 308-310 Metcalfe, Williams, Castka, pp. 335-357 Smoot, Price, Barrett, pp. 227-228	<u>Northeastern Pennsylvania Audio-Visual Center</u> EQUILIBRIUM, #A-802, (24 min. - Color) <u>McGraw-Hill Film</u>, Set #5 EQUILIBRIUM, #109210, #4124 The film deals with three questions: What is chemical equilibrium? How does the chemist recognize it? How does he explain it? In answering the questions, the film stresses the dynamic nature of equilibrium. Radioactive iodine tracers are used to demonstrate the dynamic molecular behavior of the substances at equilibrium in a closed system. An analogy in terms of fish population in two connected models present the concepts with striking simplicity. (24 min. - Color) <u>Yale Film Loop (Chem Series)</u> DECOMPOSITION OF HgO, #YF-216, (3 min. - Color) Based on the reversible reaction in a closed system.

EQUILIBRIUM

CONTENT	LEARNING EXPERIENCES
B. Equilibrium Constants 1. Law of Mass Action 2. Relationship of reaction equation to equilibrium expression	Derivation of the Equilibrium Constant Equation: $aA + bB \rightleftharpoons cC + dD$ <u>From Law of Mass Action:</u> $S_1 \propto [A]^a \times [B]^b$ $S_1 = K \times [A]^a \times [B]^b$ $S_2 \propto [C]^c \times [D]^d$ $S_2 = K_2 \times [C]^c \times [D]^d$ $S_1 = S_2$ $K_1 \times [A]^a \times [B]^b = K_2 \times [C]^c \times [D]^d$ (Products over reactants in U.S.) $\frac{K_1}{K_2} = \frac{[C]^c \times [D]^d}{[A]^a \times [B]^b}$ $\frac{K_1}{K_2} = K_{eq}$ $K_{eq} = \frac{[C]^c \times [D]^d}{[A]^a \times [B]^b}$
TEXT REFERENCES	AUDIO-VISUAL AIDS
Baker, Bradbury, et al., pp. 320-324 Metcalfe, Williams, Castka, pp. 356-359 Smoot, Price, Barrett, pp. 348-352	<u>FOM Filmstrip</u> THE EQUILIBRIUM CONSTANT, #658 Physical and chemical equilibria investigated as dynamic phenomena. Derivation of equilibrium constants for homogeneous and heterogeneous systems including Ionization Constants and Solubility Product Constants. <u>McGraw-Hill, Set # 6</u> EQUILIBRIUM CONSTANT-Keq, #405490

EQUILIBRIUM

CONTENT	LEARNING EXPERIENCES
C. Factors Affecting Equilibrium 1. LeChatelier's Principle 2. Temperature 3. Pressure 4. Concentration 5. Common Ion Effect	The factors which affect equilibrium can be discussed quite easily with the usual equation: $\text{N}_2 + 3\text{H}_2 \rightarrow 2\text{NH}_3$ <u>LAB EXPERIMENT:</u> <u>"Equilibrium", Exp. 29, p. 221, Chemistry Workbook, Dull, Metcalfe, and Williams. (Holt)</u>
TEXT REFERENCES	AUDIO-VISUAL AIDS
Baker, Bradbury, et al., pp. 310-314 Metcalfe, Williams, Castka, pp. 163-164 360-362 Smoot, Price, Barrett, pp. 352-354	<u>FOM Filmstrip</u> <u>LECHATelier's PRINCIPLE, #6502</u> Explores the operation of LeChatelier's Principle in physical equilibria. Presents the Law of Chemical Equilibrium and the role of the principle as it operates under varying sets of conditions. <u>Yale Film Loop (Chem Series)</u> <u>LECHATelier's PRINCIPLE, #YF-225</u> Relates H_2O-ice equilibrium to the pressure of a weighted-wire cutting through a block of ice. (3 min. - Color) <u>Coronet Films</u> <u>IONIC EQUILIBRIUM</u> A preparation of cupric bromide solution is used to demonstrate several features of ionic equilibrium. The "common ion" effect is discussed and demonstrated by using acetic acid and a soluble acetate salt, and hydrochloric acid and a soluble chloride salt, along with appropriate indicators. The formation and solution of precipitates is considered. (16 min. - B/W)

EQUILIBRIUM

CONTENT	LEARNING EXPERIENCES														
D. Composition of a Chemical Reaction 1. By liberation of a gas 2. By formation of a precipitate 3. By formation of an unionized product	Demonstrate precipitation and/or liberation of a gas using Chem Ed. Tested Demonstration, Ionic Reactions in Slow Motion, Dutton, <u>Tested Demonstrations in Chemistry</u>, 6th edition, p. 202.														
TEXT REFERENCES	AUDIO-VISUAL AIDS														
Baker, Bradbury, et al., pp. 318-319 Metcalf, Williams, Castka, pp. 363-364 Smoot, Price, Barrett, pp. 355-356	<u>FOM Filmstrip</u> REVERSIBLE REACTIONS AND REACTIONS TO COMPLETION, #6503 Concepts of exothermic and endothermic reactions, activation energy, conditions determining the direction of reversible reactions, and approach to equilibrium in closed systems explored. Reaching completion because of volatility, precipitation, and neutralization illustrated and explained. <u>Yale Film Loop</u> ION REMOVED BY METATHESIS, #YF-236 Utilizes neutralization of $\text{Ba}(\text{OH})_2$ by H_2SO_4. (4 min. - Color) <u>PSP Transparencies</u>, Set #TC-4 REVERSIBLE REACTIONS AND REACTIONS TO COMPLETION Illustrates by means of graphs and experiments the conditions under which a reversible reaction will reach equilibrium or go to completion. Presents equilibrium as a dynamic state that can occur in a closed system. Gives energy consideration, involved in reversing a chemical reaction, and the factors that cause a reaction to go to completion. <table data-bbox="852 2198 1850 2553"> <thead> <tr> <th><u>Transparency Titles</u></th><th><u>Overlays</u></th></tr> </thead> <tbody> <tr> <td>Reactions: To Equilibrium; To Completion</td><td>2</td></tr> <tr> <td>A Reversible Reaction: To Equilibrium; To Completion</td><td>2</td></tr> <tr> <td>Energy and a Reversible Reaction</td><td>2</td></tr> <tr> <td>Reactions to Completion</td><td>2</td></tr> <tr> <td>A Closed System Reaching Equilibrium</td><td>2</td></tr> <tr> <td>(Set of 5 Transparencies, 10 Overlays)</td><td></td></tr> </tbody> </table>	<u>Transparency Titles</u>	<u>Overlays</u>	Reactions: To Equilibrium; To Completion	2	A Reversible Reaction: To Equilibrium; To Completion	2	Energy and a Reversible Reaction	2	Reactions to Completion	2	A Closed System Reaching Equilibrium	2	(Set of 5 Transparencies, 10 Overlays)	
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NUCLEAR CHEMISTRY

THE APPROACH

The unit may be introduced with a discussion or demonstration of the classical Becquerel experiment, in which photographic plates in a light-tight box were fogged by unseen radiations from a uranium source placed atop it. The work of the Curies, Rutherford, and others should be mentioned in the solution of this atomic mystery--radioactivity.

The fact that certain combinations of A, Z, and N--the magic number rules--play a part in the prediction of the stability of a nuclide leads credence to the belief of energy levels within the nucleus itself should be stressed. Radioactivity may also be related to the major conceptual scheme that systems tend to go from higher energy levels to lower energy levels. The GE Chart of the Nuclides can be used to show most of the known nuclides, either natural or artificial.

The SMIP Filmstrip, "Radioactivity", in parts one and two, should be shown before discussing the properties of radiations from radio nuclides. This open-ended filmstrip can serve as a substitute for a lab experiment of the same type, especially if the school has little or no funds for the purchase of radiation detection instruments.

The concept of half-life should be covered thoroughly. If the students have the necessary mathematical background, they could be assigned half-life problems using a simplified form of the decay equation:

$$t_{\frac{1}{2}} = \frac{0.301 (t_2 - t_1)}{\log N_1 - \log N_2}$$

In either case, the student should be able to determine the half-life graphically if given semi-log paper and the necessary data. The half-life of a radioactive substance may be demonstrated using a radioisotope generator. (See SMIP #CC-216 for details of operation of such generators). For reasonable cost, the teacher now has available a constant source of radioactive material of half-life so short that the concept may be demonstrated in one class period. With a radioisotope generator, and a Civil Defense V-700 Geiger counter, any school should be able to perform a half-life demonstration. Some county Civil Defense directors are quite willing to lend the school enough V-700 units to enable the school to have a laboratory experiment instead of a demonstration. (See SMIP #CC-212 for details on adapting the CD V-700 counter for high school use, including simple lab experiments).

The concepts of fission, fusion, and nuclear energy should be discussed and explored in depth. The AEC and commercial organizations have

many overhead transparencies which illustrate these concepts quite well. The fact that nucleons yield so much energy -- 10^6 times more -- than ordinary chemical reactions should also be stressed.

An experiment which can be used in all schools, even those with limited facilities, is the separation of thorium-234 from uranium-238 by means of paper radiochromatography, with a 50-50 volume mixture of benzene and acetone as the developer. (See SMIP #CC-209 for the experimental procedure). This experiment also reinforces the concept of isotopes.

The last section on applications of radioisotopes may be treated adequately either with one of the AEC films or with transparencies made from the AEC booklet, "Illustrations of Radioisotopes" (Free in single copies from USAEC, Div. of Technical Information, Oak Ridge, Tenn. 37830).

SMIP has recently inaugurated a radioisotope service for teachers of Region M. Teachers may borrow Ga-68, Ba-137, and In-113 radioisotope generators at no cost from the SMIP office. License-free quantities of P-32 may also be secured free, if two-weeks prior notice is given.

OBJECTIVES

1. The student should be able to place the discovery of radioactivity in its historical perspective in the development of the atomic theory.
2. Given a typical nuclide, the student should be able to predict whether or not it will be radioactive using the instability rules. (Borderline cases -- those which obey one rule and do not obey others -- should not be assigned).
3. The student should be able to recall the properties of the four major radiations emitted from radionuclides.
The student should be able to write the symbols, including A, Z, and N, for all the major radiation types.
4. The student should be able to relate the atomic number, neutron number, and mass number of a nuclide, given its chemical symbol and mass number.
5. The student should be able to compute the half-life of a radioisotope graphically, given the original activity, final activity, and elapsed time.
6. The student should be able to recall the principles of operation of the four radiation detection instruments given.
7. The student should be able to balance nine out of ten nuclear reactions of average difficulty.
8. The student should be able to differentiate between fission and fusion reactions.
9. Given the necessary data, the student should be able to compute the energy release in MeV or calories from a balanced nuclear reaction.

10. The student should be able to name three examples of radioisotope applications from these categories:
 - a. power
 - b. radiation properties
 - c. tracers
11. The student should be able to write a short essay in favor of peacetime uses and benefits of atomic energy.

NUCLEAR CHEMISTRY

CONTENT	LEARNING EXPERIENCES
A. Natural Radioactivity 1. Discovery of radioactivity 2. Nature of radioactivity a. Instability of nuclei i. Neutron-proton ratio ii. Stability diagonal iii. Magic number rules	A-1 Class report on life of Madame Curie A-2 Repeat Becquerel experiment with sheet film in light-tight envelope and uranium ore. A-2a Demonstrate GE "Chart of Nuclides" and show how stable nuclides stay within the stability diagonal.
TEXT REFERENCES	AUDIO-VISUAL AIDS
Baker, Bradbury, et al., pp. 413-420 Metcalf, Williams, Castka, pp. 525-527 Smoot, Price, Barrett, pp. 557, 563-566 <u>Sourcebook on Atomic Energy</u>, S. Glasstone; D. VanNostrand, 1967, 3rd Ed. This book is considered the fundamental textual reference in radioactivity. <u>The Atom: A Simplified Description</u>, G. Bush, A. Silvidi; Barnes & Noble, 1961. (Good reference with little mathematical treatment). <u>Nuclei and Radioactivity</u>, G. Choppin; W. A. Benjamin, 1964. (Extensive mathematical treatment for teacher reference).	<u>McGraw-Hill Film</u> RADIOACTIVITY, #612016 A description of the basic characteristics of alpha, beta, and gamma radiation and the instruments and methods by which they may be detected and measured in terms of charge and mass. Attention is then focused on the nucleus of the atom as the source of radiation. (Color - 12 1/2 min.) <u>Atomic Energy Commission</u> INVISIBLE BULLETS This film introduces the basic knowledge about radiation necessary for an understanding of the other films in the series. The meaning of radiation, its natural sources, the various forms it takes, and how it is used in research are explained. The difference between alpha and beta particles and between gamma rays and X-rays is described. [(1962) 29 min. B & W]

NUCLEAR CHEMISTRY

CONTENT	LEARNING EXPERIENCES
A-2 (Continued) b. Radiations from nuclei - Alpha rays and properties - Beta rays and properties - Gamma rays and properties - Neutrons and properties - Nomenclature of nuclides B. Radioactive Decay - Decay constant - Half-life concept - Decay chains	2b(i). Using Geiger counter, show background radiation; absorption of α, β, γ rays by air, cardboard, lead, aluminum; effect of distance upon radiation; deflection of β particles by magnetic field. 2b(ii). Assign class drill in nuclide nomenclature. B-2 Demonstrate half-life with radioisotope generators, or milk radioisotope generator and allow class determine half-life graphically. B-3 Separate U-238 and Th-234 by means of paper chromatography.
TEXT REFERENCES	AUDIO-VISUAL AIDS
Baker, Bradbury, et al., pp. 421-426 Metcalfe, Williams, Castka, pp. 527-531 Smoot, Price, Barrett, pp. 559-563 <u>Secrets of the Nucleus</u>, J. S. Levinger, NSTA Vistas of Science Series, NSTA, 1967. (Excellent treatment, including student activities). <u>Experiments in Nuclear Science</u>, Chase et al.; Burgess, 1964. Elementary introductory and advanced laboratory experiments with simple mathematical treatment. Comprehensive teacher's guide available.	<u>Atomic Energy Commission</u> MAN AND RADIATION This popular-level film, suitable for audiences from junior high school through college, discusses many aspects of radiation and offers a survey of their widespread beneficial applications in medicine, industry, agriculture, power and research. A historical survey of the discovery of radiation is followed by an animated explanation of different types of radiation, including alpha, beta, and gamma. A brief explanation of radioisotopes and how they are produced is given, followed by scenes depicting some of their uses, including the use of Calcium-47 to diagnose bone cancer. The detection and study of radiation by sensitive instruments is explained. [(1963) Color - 28 min.] <u>McGraw-Hill Film</u> RADIOACTIVE SERIES, U-238, #626604 The film traces the various stages in the decay of U-238 to stable lead. Alpha emission, beta emission and the statistical nature of the process are emphasized. A brief mention is made of other radioactive series. (9 min. B & W)

NUCLEAR CHEMISTRY

CONTENT	LEARNING EXPERIENCES
C. Radiation Detection 1. Film 2. Electroscope 3. Cloud chamber 4. Geiger counters D. Artificial Radioactivity 1. Nuclear Reactions a. Transmutation b. Conventions in writing nuclear equations 2. Nuclear Energy a. Fission b. Fusion c. Energy from nuclear reaction (Einstein equation)	C-1 Demonstrate how radiations will fog film; demonstrate dental X-ray film packet. C-2 Demonstrate discharge of simple electroscope by radioactive source. C-3 Demonstrate tracks of radiations in diffusion cloud chamber. C-4 Demonstrate Geiger tubes and Geiger plateau of detector. D-1 Drill on nuclear equations. D-2 Work problems on Einstein equation; show tremendous energy release.
TEXT REFERENCES	AUDIO-VISUAL AIDS
Baker, Bradbury, et al., pp. 428-435 Metcalf, Williams, Castka, pp. 527-539 Smoot, Price, Barrett, pp. 568-572 <u>U. S. Atomic Energy Commission</u> <u>Understanding the Atom Series</u> Nuclear Clocks Atoms, Nature, and Man Atoms at the Science Fair Nuclear Terms, A Brief Glossary Research Reactors Synthetic Transuranium Elements	<u>McGraw-Hill Film</u> NUCLEAR REACTOR, #638002 Starting with a discussion of nuclear stability and neutron-induced fission, the film develops the ideas of neutron emission, the self-sustaining chain reaction, and the role of the moderator. It then illustrates the application of these ideas to the pile. The production of plutonium and of tracer elements is discussed. (9 min.- B & W) <u>Civil Defense</u> TALE OF TWO CITIES This film depicts the destruction of Hiroshima and Nagasaki in 1945. Although a Civil Defense film, it illustrates the tremendous energy release from a nuclear reaction in a most vivid manner. (17 min. B & W)

NUCLEAR CHEMISTRY

CONTENT	LEARNING EXPERIENCES
D (Continued) 3. Radioisotope Applications a. Power from reactors b. Properties of radiation --- gauging, etc. c. Tracers i. Tracing reaction pathways ii. Analytical uses	D-3 If time is available, an excellent experiment entitled "Detergent Efficiency" can be performed with almost any radioisotope to show how the washing efficiency of a detergent may be determined. This experiment is extremely popular with students. ("Experiments in Nuclear Science," Chase et al., p. 146)
TEXT REFERENCES	AUDIO-VISUAL AIDS
Baker, Bradbury, et al., pp. 436-458 Metcalf, Williams, Castka Smoot, Price, Barrett <u>U. S. Atomic Energy Commission</u> <u>Understanding the Atom Series</u> Power from Radioisotopes Neutron Activation Analysis Radioisotopes in Industry	<u>Atomic Energy Commission</u> RADIOISOTOPES: SAFE SERVANTS OF INDUSTRY With emphasis on safety, this film surveys the widespread uses of radioisotopes in industry. Animated explanations of the principles involved in radioisotope gauging instruments, tracing, and radiography are given. Applications of these principles are shown in various processes in the food industry, automotive research, road construction, heavy industry, oil refining and shipping, and system trouble-shooting. [(1963) Color - 28 min.] Transparencies may be made directly from the AEC publication, "Illustrations of Radioisotopes", available free from: U. S. Atomic Energy Commission Div. of Technical Information Extension Educational Materials Section P. O. Box 62 Oak Ridge, Tenn. 37831

SIGNIFICANT FIGURES

Equations and other scientific descriptions involve numbers, units, and symbols acquired from measurements of physical quantities. These numbers and units are the results of a comparison obtained from a human observer using an instrument of some type.

Certain human counting observations provide an apparent exact utilization of numbers, such as in the counting of money or people. If there are any errors made in the counting process, these errors are introduced solely by the human observer. For example, the multiplication of \$2.50 by 10 can only result in the answer \$25.00 unless the human observer has performed some error in the calculation.

However, in addition to the human factor illustrated above, scientific measurement of physical quantities involve a degree of accuracy determined by the nature and precision of the instrument used. The limitations of the instrument determine the degree of accuracy reported. Every instrument used for the measurement of length, for example, is not able to measure the quantity $\frac{1}{10,000}$ meter. Not only are instruments limited by the inherent nature of their construction, but the instruments are also limited by the human technique involved in their use.

The concept of significant figures has been derived in order to present the maximum amount of valid information and no misinformation.

Consider the following measurement: One side of a table is measured with a ruler which markings enable the observer to determine the length to be 65.78 cm. The observer measures the other side of the table with another rule and finds it to be 90.1 cm. The observer is then told to find the area of the table.

In the measurement of this table, the markings on the first rule could be estimated to the nearest hundredth of a cm -- the last digit is merely the observer's estimate that the number is closer to 65.78 than either 65.77 or 65.79. (Four significant figures.) The last figure is doubtful, but still significant. The other side of the table was found to be 90.1 cm, rather than 90.0 or 90.2 cm. (Three significant figures.) The area, upon multiplication, may be written with only three significant figures, the number of s.f. in the least reliable number.

$$\begin{array}{r}
 65.7\textcircled{8} \\
 90.\textcircled{1} \\
 \hline
 \textcircled{6}\textcircled{5}\textcircled{7}\textcircled{8} \\
 592020 \\
 \hline
 5926.778
 \end{array}$$

all these are multiplied by the doubtful figure 1, so they are all doubtful.

The answer should have 3 s.f., or should be 5930.

RULES FOR SIGNIFICANT FIGURES

In general, every digit from one through nine (1, 2, 3, ..., 9) in a given number is a significant figure.

EXAMPLES

2.5 indicates 2 s.f.

2.57 indicates 3 s.f.

The significance of the zero is dependent upon the position of the zero in the number. It is considered good practice among some workers to use the zero to locate the decimal point of a number, such as 0.1 (1 s.f.), but the zero in such a case is not a significant figure.

Zeroes are significant when positioned between two other digits, or when positioned at the end of a number.

EXAMPLES 1.02 indicates 3 s.f.
 1.200 indicates 4 s.f.

Zeroes are not significant when the first digit of a number, or when preceded by other zeroes.

EXAMPLES 010 indicates 2 s.f.
 0.010 indicates 2 s.f.

In addition or subtraction, disregard all digits to the right of the column which has the most number of significant digits.

EXAMPLES 3.5782 3.6
 1.245 1.2
 6.30 6.3
 5.2 5.2
 16.3

In multiplication, the number of significant figures in the answer should be the same as the number of significant figures in the least reliable measurement

35.681
5.92

The position of the decimal point in a number has little to do with the number of significant figures.

PERCENTAGE OF ERROR

Most of the quantities which you will measure in your experiments have already been measured with painstaking care and with great accuracy and precision. These accepted values usually may be found in a handbook or in the appendix of a text or lab manual. Many times your values will differ from these accepted values. As an indication of your technique, you will be expected to find the percentage of error of your measurement.

$$\text{Pct error} = \frac{\text{difference between observed value and accepted value} \times 100}{\text{accepted value}}$$

EXAMPLE

The index of refraction of a substance is found to be 1.33 by a student. The accepted value from a handbook is 1.35. What is the percentage of error?

$$\text{Pct error} = \frac{1.35 - 1.33}{1.35} \times 100 = 1.5\%$$

A percent error of greater than 5% indicates a serious deficiency in technique.

Occasionally, there is no accepted value to go by, or the accepted value has been obtained by means of a technique different from the one you may be using. Then the percent error is found thus:

$$\text{Pct error} = \frac{\text{difference between values}}{\text{either value}} \times 100$$

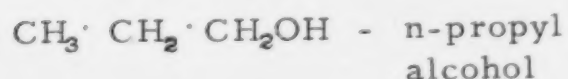
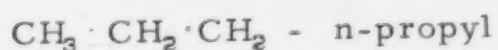
NOMENCLATURE OF ORGANIC COMPOUNDS

Organic compounds may have (1) common or "trivial" names, (2) derived names, and (3) systematic or IUC names. No one method is used exclusively. It is possible to produce several names for a specific organic compound. Ethyl alcohol (common name) is also ethanol (IUC) and could be named hydroxyethane (derived name) or methyl carbinol.

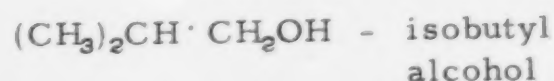
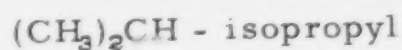
(1) Common name nomenclature developed in the past. The source of many such names is hidden. Most of these names give no hint as to the structure, so formulas for such compounds as acetone, glutaric acid, xylene, or cortisone can be written only from memory. Some names are so widely used that the systematic name is used only rarely; with others the common name is so much shorter. This is particularly true among the fused aromatic and alicyclic systems.

The terms "normal," "iso," "secondary," and "tertiary" are used with certain small hydrocarbon groups.

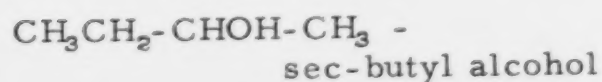
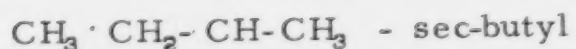
"Normal"



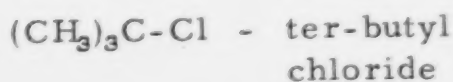
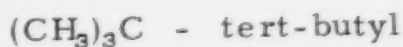
"Iso"



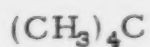
"Secondary"



"Tertiary"



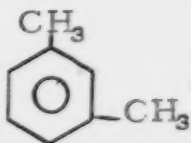
(2) The second system is useful as compounds are named as derivatives formed by substitution of hydrogen by various atoms or radicals. Some examples follow:



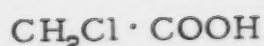
tetramethyl methane



1, 2, 3 trimethyl cyclopropane
(also the IUC name)

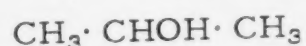


m-dimethyl benzene or IUC
1, 3-dimethyl benzene or
m-xylene



chloroacetic acid

Alcohols are also named as derivatives of methyl alcohol - carbinol being the name for methyl alcohol.



dimethyl carbinol also called
isopropyl alcohol (common)
2-propanol (IUC)

(3) Any organic compound can be named using principles developed and recognized by the International Union of Pure and Applied Chemistry, commonly called the IUC by most American chemists. Some chemists also refer to this system as "The Geneva System," from the meeting place where it was adopted.

Every IUC name may contain three parts:

1. The root of the name, which signifies the length of the carbon chain.
2. A suffix, which indicates the major functional group on the chain.
3. A prefix, which indicates the minor functional groups on the chain.

In addition, numbers may be attached to the prefixes and suffixes to identify the position of the functional groups on the chain. The user of this system must learn, (1) to spot the longest chain of carbon atoms even though the atoms are not written in a straight line, (2) the names of the hydrocarbon group or root, and (3) the prefixes and suffixes indicating the functional group or groups.

Root - the length of the longest straight chain of C atoms determines the root.

<u>C number</u>		<u>Root</u>	<u>Name of Alkane</u>
1	-C-	<u>Methyl</u>	<u>Methane</u>
2	-C-C-	<u>Ethyl</u>	<u>Ethane</u>
3	-C-C-C-	<u>Propyl</u>	<u>Propane</u>
4	C-C-C-C-	<u>Butyl</u>	<u>Butane</u>
5	C-C-C-C-C-	<u>Pentyl</u>	<u>Pentane</u>
6	C-C-C-C-C-C-	<u>Hexyl</u>	<u>Hexane</u>
7	C-C-C-C-C-C-C-	<u>Heptyl</u>	<u>Heptane</u>
8	C-C-C-C-C-C-C-C-	<u>Octyl</u>	<u>Octane</u>
9	C-C-C-C-C-C-C-C-C-	<u>Nonyl</u>	<u>Nonane</u>
10	C-C-C-C-C-C-C-C-C-C-	<u>Decyl</u>	<u>Decane</u>

The longest chain may not be readily apparent from the formula, depending upon the way the structural formula is drawn.

- a) Longest chain _____ 5 C atoms $\begin{array}{c} \text{C}-\text{C}-\text{C}-\text{C} \\ | \\ \text{C}-\text{C} \end{array}$
- b) Longest chain _____ 6 C atoms $\begin{array}{c} \text{C}-\text{C} \\ | \\ \text{C}-\text{C}-\text{C}-\text{C} \\ | \\ \text{C}-\text{C}-\text{C} \end{array}$

Remember that in the tetrahedral carbon atom, all four positions are equivalent.

- c) Longest chain _____ 3 C atoms $\begin{array}{c} \text{C} \\ | \\ \text{C}-\text{C}-\text{C} \\ | \\ \text{C} \end{array}$
in any direction

The nature and position of a functional group on the chain is indicated by a prefix for the name of the functional group and a number to indicate position of the chain.

- a) 2-chloropropane
also chlorodimethyl methane
(der. name)
 $\begin{array}{c} \text{CH}_3\text{CH}-\text{CH}_3 \\ | \\ \text{Cl} \end{array}$
- b) 2-chloropropane
(same as number 1)
 $(\text{CH}_3)_2\text{CHCl}$
- c) 1-chlorobutane
 $\text{CH}_2\text{Cl} \cdot \text{CH}_2 \cdot \text{CH}_2 \cdot \text{CH}_3$
- d) 1-chlorobutane
 $\text{CH}_3\text{CH}_2\text{CH}_2 \cdot \text{CH}_2\text{Cl}$
- e) 2-bromopentane
NOT 4-bromopentane
 $\begin{array}{c} \text{CH}_3 \cdot \text{CH}_2 \cdot \text{CH}_2 \cdot \text{CH} \cdot \text{CH}_3 \\ | \\ \text{Br} \end{array}$
- f) 2,2 dibromobutane
 $\begin{array}{c} \text{Br} \\ | \\ \text{CH}_3 \cdot \text{CH}_2 \cdot \text{C} \cdot \text{CH}_3 \\ | \\ \text{Br} \end{array}$
- g) 1,2 dibromobutane
 $\begin{array}{c} \text{CH}_2\text{Br} \cdot \text{CH} \cdot \text{CH}_2 \cdot \text{CH}_3 \\ | \\ \text{Br} \end{array}$ also $\begin{array}{c} \text{CH}_3 \cdot \text{CH}_2 \cdot \text{CH} \cdot \text{CH}_2 \\ | \quad | \\ \text{Br} \quad \text{Br} \end{array}$

HYDROCARBON SUFFIXES

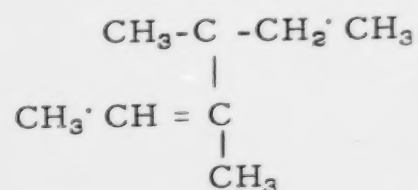
<u>Class</u>	<u>Compound</u>	<u>Suffix</u>
aliphatic (alkanes)	ethane	-ane
olefin(alkenes)	ethene	-ene
acetylene(alkynes)	ethyne	-yne
aromatic	benzene	-ene
cyclo paraffin (cycloalkane)	cyclobutane	-ane
cyclo olefin	cyclobutene	-ene

STRAIGHT CHAIN UNSATURATES

The position of double or triple bonds is indicated by adding the suffix -ene or -yne followed by the lowest number to indicate the position of the double or triple bond. (Triple bond, suffix -yne).

- a) butene - 2, 2-butene
NOT butene - 3
sym. dimethylethylene
 $\text{CH}_3 \cdot \text{CH} = \text{CH} \cdot \text{CH}_3$

- b) 3, 4-dimethylhexene - 2



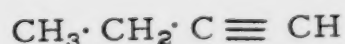
- c) butene - 1, 1-butene
NOT butene - 3
NOT butene - 4



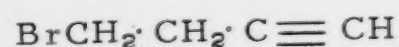
- d) heptyne - 2, 2-heptyne
NOT heptyne - 5



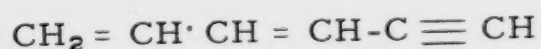
- e) 1-butyne
NOT 3-butyne



- f) 4 bromo-1-butyne



- g) 1, 3 hexadiene - 5 - yne



Chemical abstracts usually gives the lower number to the double bond.

SATURATED AND UNSATURATED ALICYCLICS

These compounds are named the same as the straight chain hydrocarbons with the prefix cyclo.

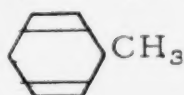
a) cyclopropane



b) cyclopentene



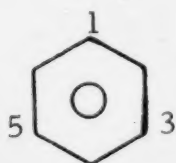
c) 3-methyl-1, 4-cyclohexadiene



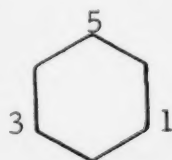
AROMATIC HYDROCARBONS

The six carbons in benzene are equivalent so a mono substituent is on position one. With two or more substituents, the numbering keeps as low as possible although ortho, meta, and para may be used for [1, 2], [1, 3] and [1, 4] positions.

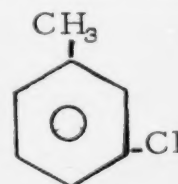
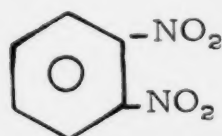
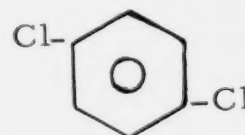
benzene



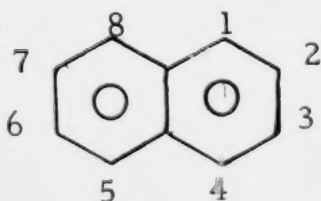
cyclohexane



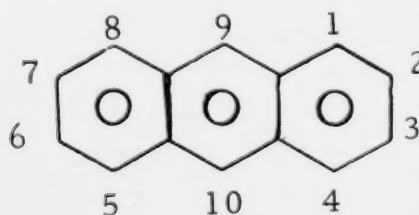
3-chlorotoluene

1, 2-dinitrobenzene
or
orthodinitrobenzene1, 4-dichlorobenzene
or
paradichlorobenzene

naphthalene



anthracene

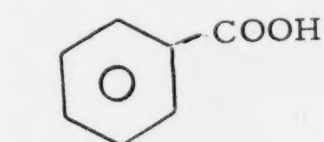


When naming multifunction compounds, both prefixes and suffixes are used. An order of precedence has evolved which serves to indicate which function will be given by the suffix. Usually only one suffix is used, but this rule is by no means absolute. The radical $\text{CH}\equiv\text{C}\cdot\text{CH}:\text{CH}\cdot\text{CH}_2^-$ would be named 2-penten-4ynyl.

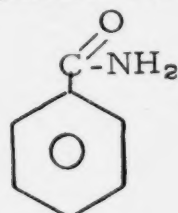
The order of precedence and the suffix indicating the functional group are given in the order of decreasing importance.

Function	Formula	Suffix
(1) acid	-COOH	-oic acid
(2) acid chloride	-COCl	-oyl chloride
(3) amide	-CONH ₂	-oic amide
(4) aldehyde	-CHO	-al
(5) nitrile	-C≡N	-ontrile
(6) ketone	>C=O	-one
(7) alcohol	-CH ₂ OH (primary) >CHOH (secondary) >COH (tertiary)	-ol
(8) phenol	-OH	-ol
(9) amine	-NH ₂ (primary) >NH (secondary) >N (tertiary)	-amine

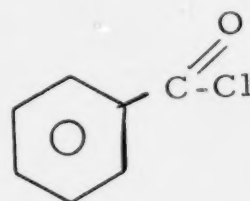
Many trivial names are used with a shift to -oyl or -yl for the acid halide, addition of -amide and -ate for the ester.



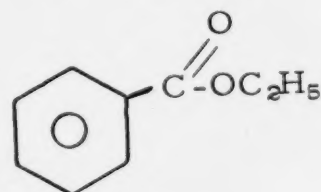
benzoic acid



benzamide

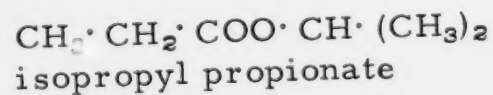
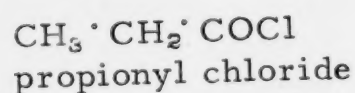
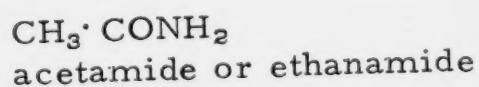
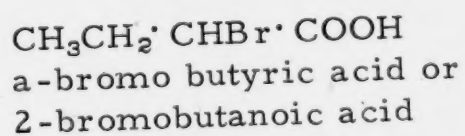


benzoyl chloride



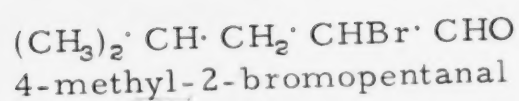
ethyl benzoate

Alpha, beta, and gamma are also used to indicate the position of a substituent. The alpha position is on the carbon next to the primary functional group.

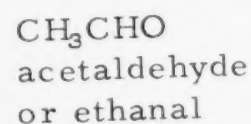
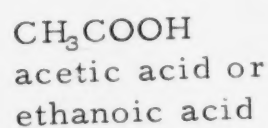


(1), (2), (3) carboxylic acid, acid halide, amide, ester
-COOH -COCl -CONH₂ -COOR

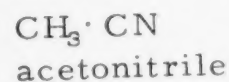
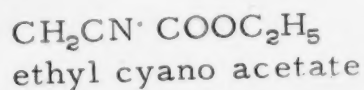
(4) -CHO is always the terminal group so the number 1 is not used--



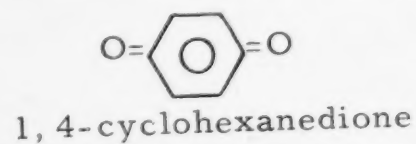
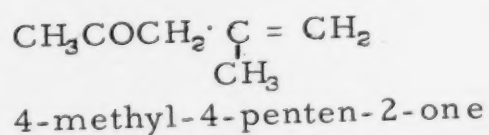
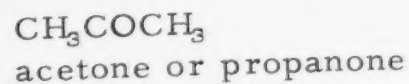
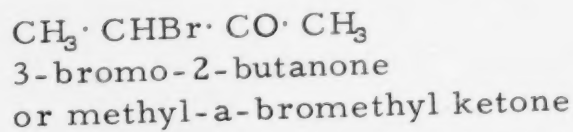
Aldehydes may also be named from the corresponding acid so the above is also a substituted valeric aldehyde or valeraldehyde.



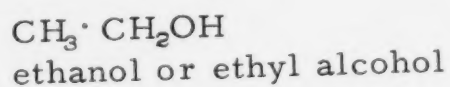
(5) $-\text{C}\equiv\text{N}$, nitrile: $-\text{C}\equiv\text{N}$ may also appear as cyano when it is outranked in a multifunction compound.



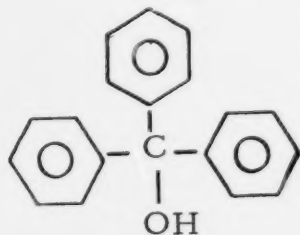
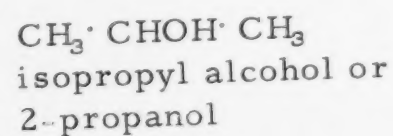
(6) -one, ketone:



(7) -ol (p) $-\text{CH}_2\text{OH}$ (s) CHOH

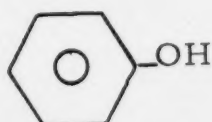


(t) COH alcohols

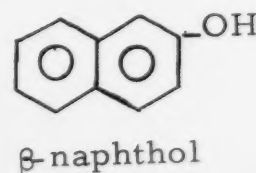
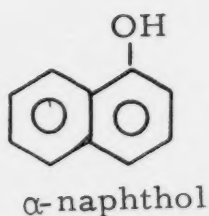


triphenyl carbinol or
triphenyl methanol

(8) $-\text{OH}$ attached to an aromatic radical:



phenol rather than
hydroxybenzene

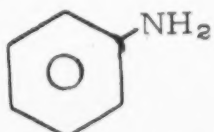


(9) -NH_2 , NH , N

primary, secondary and
tertiary amines

$\text{CH}_3 \cdot \underset{\text{NH}_2}{\text{CH}} \cdot \text{CH}_3$
isopropylamine

$(\text{CH}_3)_3\text{N}$
trimethyl amine



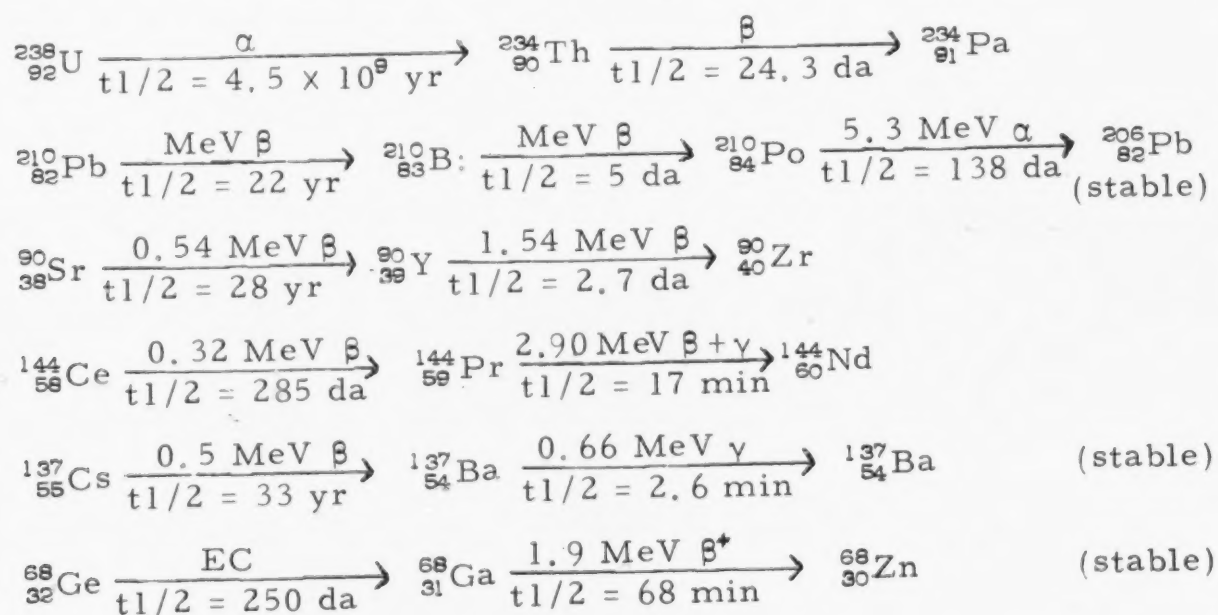
phenylamine, but better
known as aniline

The Chemical Abstracts Service, 1155 16th St., NW, Washington, D. C., 20036, publishes The Naming and Indexing of Chemical Compounds from Chemical Abstracts. This gives a thorough run down on organic and inorganic nomenclature as used in the U. S. A. today.

RADIOISOTOPE GENERATORS

Over the past decade, many nuclear scientists have explored the problem of rapidly and easily separating the short half-life daughter radioisotopes from their longer half-life parent radionuclides. Workers at Brookhaven National Laboratory and other institutions have applied ion-exchange separation techniques to this problem with great success. The fruits of their early labors are now available commercially, called either "radioisotope generators" or "radioisotope cows."

A great many radioactive decay chains are found in both natural and artificially-produced radioisotopes; in a decay chain, the radioactive decay of the original nuclide does not result in a stable nuclide but another radionuclide, called the daughter. Some familiar decay chains are:



When the parent nuclide has a half-life of 10^4 or more times greater than the half-life of the daughter, the condition is known as secular equilibrium. All the decay chains cited above are secular equilibrium chains.

A secular equilibrium chain in a sense offers an almost inexhaustible supply of short half-life material; once the mixture has been milked of the daughter, the parent will grow more daughter at the same rate at which the daughter decays. For example, in the ${}^{137}\text{Cs}/{}^{137}\text{Ba}$ decay chain, in 2.7 min. the ${}^{137}\text{Cs}$ will grow in half the amount of daughter which was removed. Once a radioisotope generator has been milked, it will grow in more daughter. The useful life of such a generator is limited then only by the half-life of the parent, rather than by the number of times daughter material has been extracted.

The generators are usually sold commercially as some form of vertical cylinder with an inlet at the top end and an outlet at the bottom end. The parent radioisotope is retained on an ion-exchange resin bed effectively sealed within the cylinder. To milk the generators, an eluting solution—of type dependent upon the parent/daughter mixture—is added to the inlet orifice; in a few minutes or less the eluant flow through, carrying the released daughter along with it through the outlet tube. Eight half-lives of the daughter later, the generator is at full charge of daughter again and may be milked once more.

Thus, the radioisotope generator is an educational tool with few peers—it enables the science teacher to have short-lived radioactive materials available in the laboratory at a moment's notice. It is no longer necessary to order short-lived radioisotope two weeks in advance many times a year; one purchase and a supply of radioactive material is constantly available in the laboratory.

Radioisotope generators are available from these companies at this writing:

New England Nuclear	General Radioisotope	Union Carbide
*Ga-68 (68 min)	Y-90 (2.7 da)	*Ba-137 (2.7 min)
	Pr-144 (17 min)	*In-113 (104 min)
	Ba-137 (2.7 min)	
	Ga-68 (68 min)	

(Half-life of daughter in parentheses.)

Quantities of parent isotope supplied by the manufacturers differ; some require AEC specific licenses, while others are generally-licensed or "license-free."

The above generators have a number of uses in any educational program. Using Ba-137 with its extremely short 2.7 min half-life, the concept of half-life can be demonstrated or investigated in the lab in one short lab period. In-113 ($t_{1/2} = 104$ min) or Ga-68 ($t_{1/2} = 68$ min) can also be used for lab investigations to show half-life and to perform simple qualitative tracer experiments.

For quantitative experiments in biology, notably uptake of P^{32} in frogs, mice, etc., and autonadiography in plants, a radioisotope with a longer half-life is more desirable, such as $P-32$ with its 14.3 da. half-life. All of these isotopes emit radiations with energies able to be detected with Civil Defense V-700 or Classmaster Geiger Counters found in most high

schools. SMIP also has filmstrips available for loan. For a complete unit on radioactivity, SMIP offers this package on a loan basis:

Radioisotope generators
Teacher Curriculum Materials
Nuclear Technique Filmstrips
Constant-Geometry Tube Stand for CD V-700 Counters

SMIP offers to the science teacher of Region M an opportunity to borrow those generators marked on the previous page with an asterisk. Call or write the SMIP office to reserve your generator. The only conditions are these:

- (1) Please call beforehand.
- (2) Please carefully follow instructions for milking packed with each generator.
- (3) Please pick up generators at the SMIP office.
- (4) Please return generators promptly after completion of use.

CALL OR WRITE

SMIP RADIOISOTOPE SERVICE
WILKES COLLEGE
BOX 111
30 WEST RIVER STREET
WILKES-BARRE, PENNSYLVANIA 18703

PHONES: DAY 824-4651 Ext. 305
NIGHT 825-6932

GAS VOLUME AND PRESSURE

Problem:

To determine the effect of pressure on the volume of a gas.

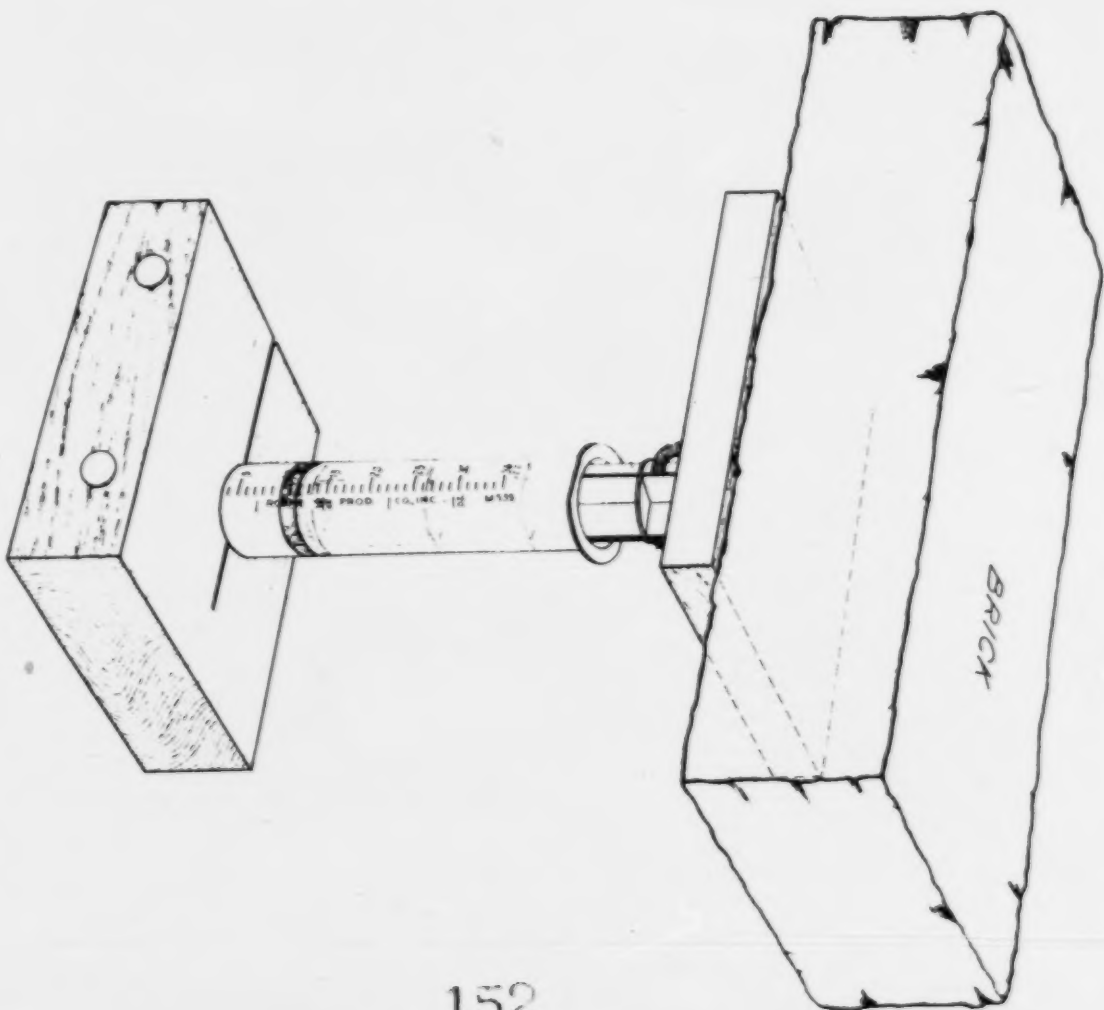
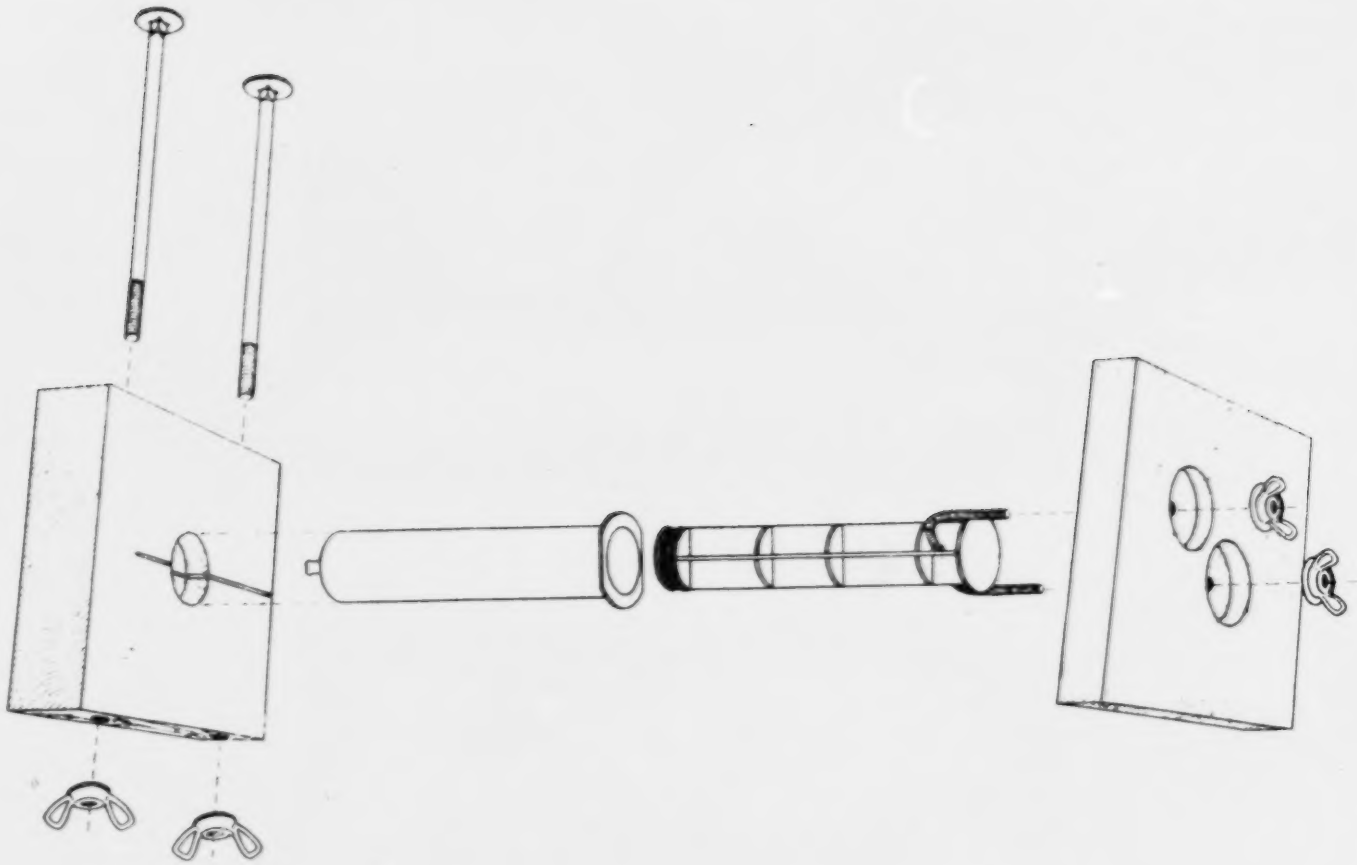
Introduction:

In this experiment, you will increase the pressure on a volume of air enclosed in a cylinder. In a right circular cylinder, the volume is directly proportional to the height, since the base diameter is kept constant. The changes in volume will be indicated by the changes in the markings on the syringe used as the cylinder. The pressure will be due to the bricks added to the cylinder piston. If the bricks are identical, our pressure units may be taken as the number of bricks added.

Materials:

IPS Elasticity of Gases apparatus

Four or five typical masonry bricks



EDUCATIONAL DIVISION, DAMON ENGINEERING, INC.

NEEDHAM HTS., MASS.

Procedure:

1. Insert the "U" bolt into the plunger as shown in figure A. Affix the smaller wing nut to the "U" bolt and tighten.
2. Insert the syringe barrel into the base of the apparatus - the larger block - and try to squeeze the flat end of the syringe barrel into the tiny slot in the base. Insert the carriage bolts into the base, add the larger wing nuts, and tighten securely.
3. Place the thin wire into the empty syringe barrel past the graduations, while holding the other end of the wire in your right hand. With the left hand insert the plunger assembly into the syringe barrel to a depth just above the graduations. You will hear a slight hissing noise as air is flushed from the apparatus.
4. Remove the wire with the right hand, making certain the plunger is held rigid with the left hand. The apparatus is now ready for the experiment.
5. Record the reading on the graduations made by the end of the plunger.
6. Add one brick to the top of the plunger assembly. Try to balance it as well as you can. Record the reading with one brick in place.
7. Add another brick and record. Repeat until a maximum of five bricks rests on the plunger assembly. (What may happen if too many bricks are added?) Record all data carefully.
8. At the end of the experiment, dismantle the apparatus and return it to the box in which it was supplied.

Conclusions:

Draw a graph from the data, with pressure (as number of bricks) on the horizontal axis and volume on the vertical axis.

What type of relationship exists between the volume and pressure of a gas? Write a general equation to express this relationship.

DIFFUSION

PROBLEM

Find the density of a gas using Graham's Law of Diffusion.

INTRODUCTION

Diffusion may be defined as the movement of molecules from a region of high concentration to a region of low concentration. From common experience with perfumes and other odors, we know that all substances, notably gases, do not diffuse at the same rate. In this experiment, you will measure the relative rates of diffusion of two gases and use this information to calculate the density of one of the gases.

APPARATUS

Beakers, 100 and 150 ml; combustion tube, pyrex; burette clamps; metric ruler; ring stands.

MATERIALS

Conc. hydrochloric acid; conc. ammonium hydroxide; 2 straight pins; 2 corks to fit combustion tube; cotton balls; piece of black paper; masking tape; acetone.

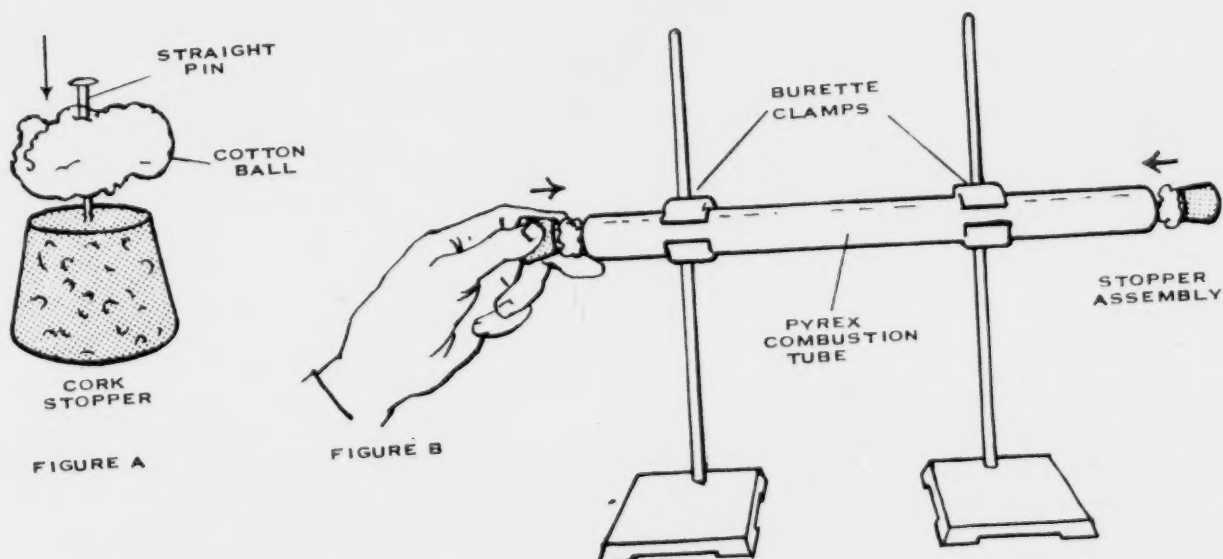


FIGURE 37

PROCEDURE

Clean a glass combustion tube with detergent and water. Rinse with distilled water. Add 10 ml of acetone to the tube and swirl the acetone so it reaches every inside surface of the tube. Discard the acetone and allow the vapors to escape. The tube should be completely dry in a few minutes, with no faint smell of acetone.

While the tube dries, prepare two cork assemblies as shown in Figure A.

Clamp the combustion tube as shown in Figure B. Place a short strip of masking tape at each end. Insert the cork assemblies into each end of the tube and mark the furthest extent of their travel on the tape. Tape a piece of black paper on the tube near its center. This will enable you to see more readily the white ring of ammonium chloride formed when the gases meet after diffusing from each end of the tube.

Pour a few ml of hydrochloric acid in one beaker and a few ml of ammonium hydroxide in another. Dip the cotton ball of the cork assembly in the HCl solution and the other in the NH_4OH solution at the same time. Now quickly insert the cork assemblies into the ends of the tubes at the same time and up to the marks. Try not to inhale any of the fumes which have been given off. The two gases now diffuse through the tube and form a white ring where they meet. Do not become impatient. Mark this spot as soon as you see it. Measure the distance to a tenth of a centimeter that each gas traveled. Record these distances in your data table.

Clean and dry the tube as before, and make two more determinations.

DATA

GAS	DENSITY	DISTANCE			
		1	2	3	avg.
Ammonia					
Hydrogen chloride	1.64 g/liter				

CALCULATIONS

Graham's Law of Diffusion may be written as follows:

$$\frac{R_1}{R_2} = \sqrt{\frac{D_2}{D_1}}$$

where R = rate of diffusion of the two gases

D = densities of the two gases

or,

$$\frac{R_1^2}{R_2^2} = \frac{D_2}{D_1}$$

Since the two gases travelled their distances in the same time, the ratio of the distances may be taken as their rate of diffusion and substituted for R_1 and R_2 in the equation. The density of hydrogen chloride and the ratio R_1/R_2 is known, leaving only the density of ammonia as the unknown. Compute the density of ammonia.

QUESTIONS

1. Why must the tube be clean and absolutely dry before using it?
2. How many times faster does one gas diffuse, compared to the other?

GAS VOLUME AND TEMPERATURE

PROBLEM

To find a relationship between the volume of a gas and its temperature.

INTRODUCTION

In this experiment you will endeavor to find a mathematical relationship between the volume of a gas and its temperature. You will heat a volume of air in a tube of uniform bore. The volume of gas will be kept in the tube by a small disk of mercury.

The tube you will use is actually a right circular cylinder. Since the volume of such a cylinder is directly proportional to its height, we may use the height of the cylinder (as indicated by the position of the mercury disk inside) instead of the volume for our calculations. Remember that π is constant along with the bore of the cylinder.

APPARATUS

Berzelius beaker, 250 ml; burette clamp; condenser clamp; extension clamp; ruler, 50 cm, metric scale; special capillary tube; thermometer, 110°C ; burner; stirring rod; ring stand and ring; wire gauze.

MATERIALS

Mineral oil; cleanser; 2 small rubber bands.

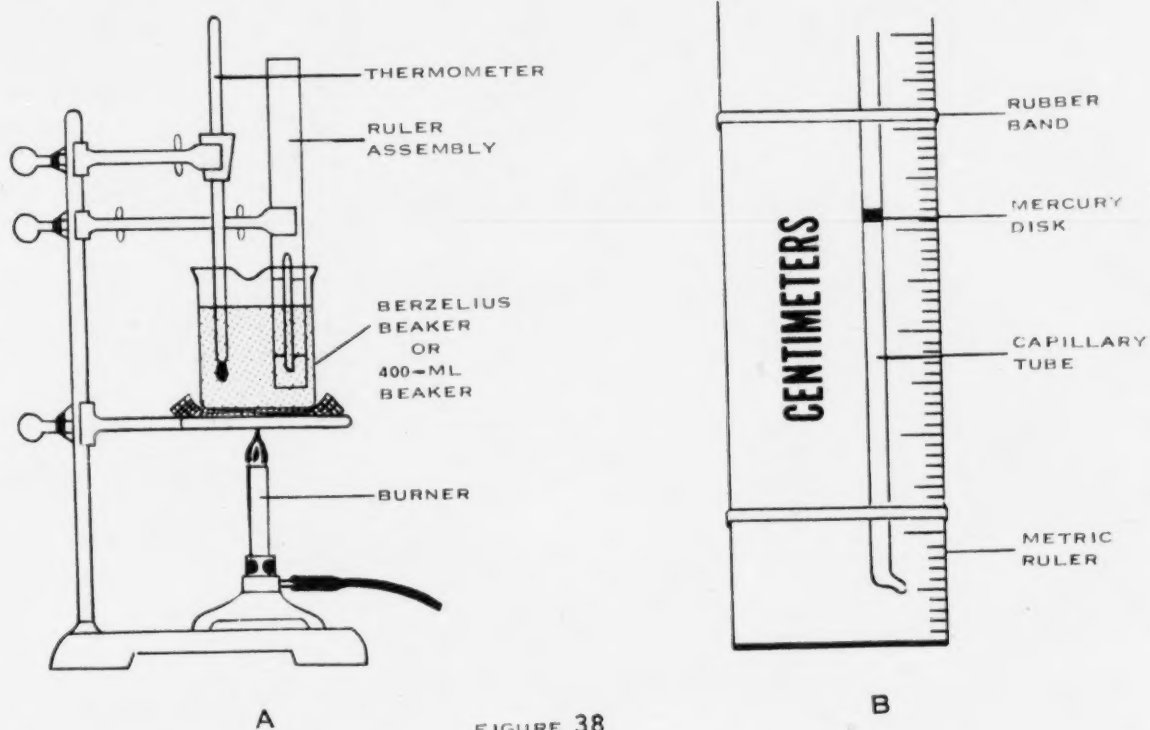


FIGURE 38

PROCEDURE

Set up the apparatus as shown in the diagram. Pour enough mineral oil into the beaker so that one inch of space remains between the top of the beaker and the level of liquid. This allows for expansion of the mineral oil. Check to see if the length of the capillary tube lies parallel with the edge of the ruler, and that the thermometer bulb is approximately even with the middle of the gas volume in the capillary tube. Wipe the outside of the beaker if you have spilled any oil.

Start heating the beaker slowly. As the temperature increases, you will notice convection currents in the mineral oil. Insert your stirring rod in the liquid and carefully agitate the mineral oil. Be careful not to bump either the thermometer, ruler, or capillary tube.

One partner should stir the liquid and call off the temperature to the other who will read the height of the air column at that temperature and record the data.

The height of the air column should be estimated to the nearest hundredth of a cm at these temperatures: 70, 75, 80, 85, 90, 95 and 100° C. The initial height of the air column is of no importance. Remember to read the height of the air column at the bottom of the mercury disk.

After the last temperature reading (100° C) has been obtained, remove the burner and allow the apparatus to cool for a few minutes. When the temperature has decreased by 20 degrees or so, elevate the thermometer and ruler assemblies to allow them to drain. The mineral oil may be replaced; it is not a reagent. The beaker may be rinsed initially with acetone to cut the mineral oil, and then thoroughly scrubbed with cleanser (Ajax, Comet, etc.) and a brush until clean.

CALCULATIONS

First draw a graph from the data, with temperature on the horizontal axis and volume on the vertical axis.

What type of relationship exists between the volume and temperature of a gas? Write a general equation to express this relationship.

QUESTIONS

1. In what way was the pressure of the gas held constant in this experiment?
2. What do you think is the reason for not taking any temperature readings below 70° C?

SECTION I

BACKGROUND AND THEORY

NATURE OF RADIOACTIVITY

One of the first men to observe the phenomenon of radioactivity was Henri Becquerel, who found that uranium salts would fog or darken photographic films even when the films were kept in light-tight containers.

The unseen "rays" from uranium salts were later found to be a mixture of three different types of rays. Elements of atomic number greater than 82 were generally found to be naturally radioactive. The emanation of the rays was also found to be the result of changes within the nucleus of the atom itself, tending toward a more stable arrangement of the nucleus.

The rays given off by radioisotopes were found to be of three types:

- (1) alpha rays--these rays are really positive particles which are identical to helium nuclei except for their nuclear origin. They travel about 0.1 the speed of light. Because of their relatively slow speed and great charge, these particles possess great ionizing power. The alpha particles emitted from a particular nucleus are identical in energy and hence are called "mono-energetic." Alpha particles may be stopped by a thin sheet of paper.
- (2) beta rays--again, these rays are really particles, but unlike alpha particles, they are negatively charged and are identical to the electron except for their nuclear origin. Beta particles travel up to 0.9 the speed of light. Unlike alpha particles, betas are emitted with a wide range of energies, hence the variation in their speed. Betas are more penetrating than alpha particles, but may be stopped by several thicknesses of metal.
- (3) Gamma rays--are really rays, similar to x-rays but usually of shorter wavelength. Since gamma rays have no charge, they do not react with matter in the same manner as alpha and beta particles. They travel at the speed of light and have great penetrating power--a foot of lead is necessary for absorption.

Later experiments showed that the nuclei of a single species of an element were not exactly alike, as postulated by the early proponents of atomic theories. The nuclei of the same element were found to be actually mixtures of nuclei which possessed the same atomic number (no. of protons) but different mass numbers (no. of protons + no. of neutrons). The different nuclei of the same element were called isotopes.

Isotopes which give off rays are called radioisotopes. Even the naturally radioactive elements (atomic number greater than 82) were found to contain many different species of isotopes. Some species have been present from the beginning of the universe while others are constantly being made artificially by man through changes brought about in the nuclei of atoms.

ACTIVITY OF RADIOACTIVE SAMPLES

The term radioactivity refers to the rate of disintegration of the radioactive nuclei within a sample, resulting in the emission of rays. It can readily be deduced that the greater the radioactivity of the sample, the greater the rate of disintegration of the radioactive nuclei.

The unit of radioactivity most commonly used in nuclear science is the curie and its multiples thereof. By definition, one curie is equivalent to that quantity of radioactive material which disintegrates at the rate of 3.7×10^{10} disintegrations per second.

one curie	$1 \text{ c} = 3.7 \times 10^{10} \text{ dis/sec}$
one millicurie	$1 \text{ mc} = 3.7 \times 10^7 \text{ dis/sec}$
one microcurie	$1 \mu\text{c} = 3.7 \times 10^4 \text{ dis/sec}$

The relative activity may be expressed by the detecting instrument in a variety of ways--counts per minute as in the case of the famous Geiger counter, clicks per minute as in Geiger counters used by uranium prospectors, darkness of the image in a photographic film, or the rate of drift of a quartz fiber as in the case of the electroscope.

THE TRACER CONCEPT

The "tracer" technique, which is used to follow the path and location of atoms and molecules, depends upon the fact that radioactive atoms of a species will behave chemically in essentially the same manner as non-radioactive atoms of the same species. For example, radioactive phosphorus atoms will undergo the same reactions as non-radioactive phosphorus. The radioactive phosphorus, however, emits detectable rays which enable an observer to determine the path and location of the phosphorus.

THE PHOTOGRAPHIC PROCESS

In the photographic process, radiation (light rays) is allowed to impinge upon a film which is coated with a gelatine emulsion containing some type of silver halide. The light causes a chemical change in the silver halide, enabling it to be "developable". Development of the film consists of immersion in a special alkaline reducing solution. This results in the formation of a metallic silver image, dark in proportion to the amount of light to which it was exposed. To preserve the image, the film is then immersed in another solution which removes all the chemicals in the film emulsion except the metallic silver. This latter process is called fixing and makes the image on the film permanent.

The resultant image is exactly the reverse of the photographed object--what was white on the photographed object is black on the film image, and vice versa. This picture is called a negative. The negative is placed in contact with paper sensitized with a silver halide, as was the film. The paper is exposed, and the processing repeated. The resultant positive picture is exactly like the photographed object.

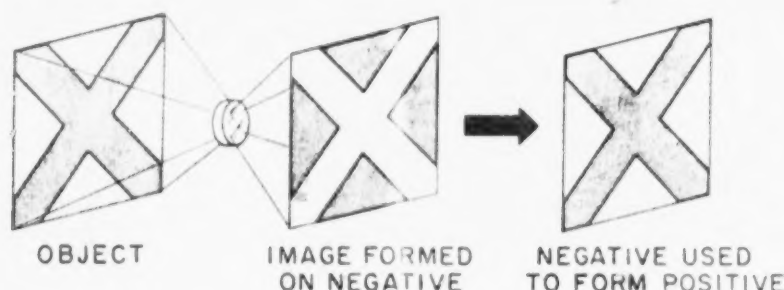


Figure 1. The Photographic Process

AUTORADIOGRAPHY--PICTURES FROM RADIATION

Toward the end of the 19th century, Becquerel found that uranium salts possessed the same ability as light to affect a photographic film. He placed a uranium sample on top of a light-tight container which held unexposed photographic plates. (On a photographic plate, the special emulsion is coated onto a glass plate rather than a film.) Upon development, the plate showed a darkening identical to the outline of the uranium sample, although no light was ever allowed to impinge upon the film! Becquerel had discovered radioactivity.

In autoradiography, the photographic process is teamed with the tracer technique to provide the scientist with a valuable tool. An object, animal, plant, or inanimate article which contains radioactive material is placed near a photographic film. The emitted radiation produces an image on the film.

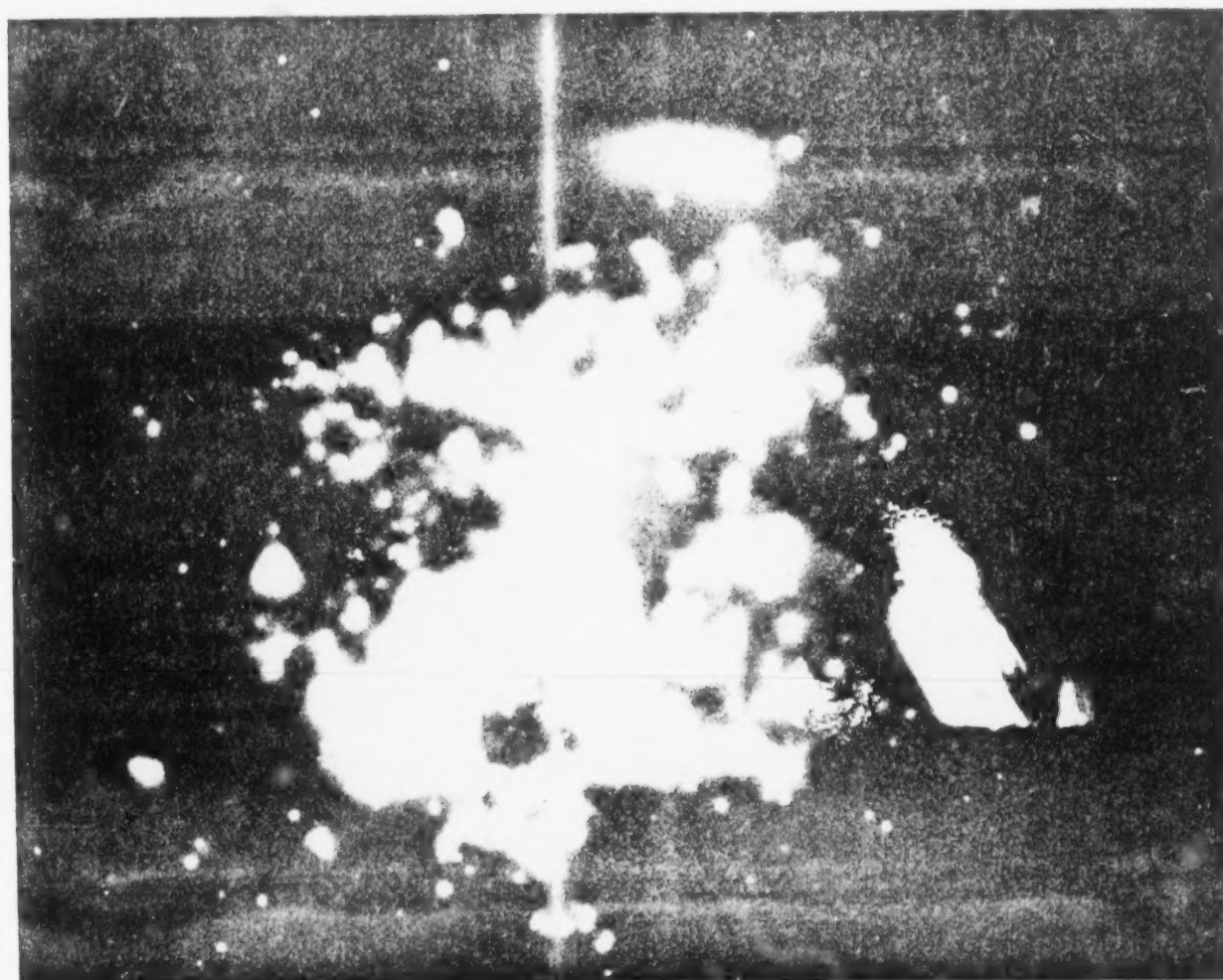


Figure 2. Autoradiogram

Not only does the radiation produce an image, but it locates the radioactive material and also gives a relative indication of the amount of radioactivity present. Another advantage is the permanent record which is secured. Probably the greatest advantage is that no electronic detection equipment is required, although such equipment is desirable in computing exposures.

PREPARING THE AUTORADIOGRAM

Essentially, there are 4 steps in making an autoradiogram. These are as follows:

- (1) Preparing the radioactive specimen;
- (2) Computing the exposure required;
- (3) Exposing the film to the radioactivity;
- (4) Processing the film.

Meticulous care must be given to every step to secure good results.

Preparing the radioactive specimen

This step will be covered in greater detail in each experiment as the technique will vary with the choice of object or organism. Plants may be grown in radioactive nutrient solutions, and objects may be plated or coated with a radioactive material or made radioactive in some other way. The specimen is placed next to the unexposed film in a completely dark room, or under light from a safelight equipped with a Wratten Series 6B safelight filter.

If the specimen is completely dry, under certain conditions, it may be placed directly next to the film. As a matter of standard practice, a sheet of thin polyethylene film such as Saranwrap is kept between the specimen and the film to prevent undesirable juices and other liquids from the sample coming into contact with the film. Today, even though ordinary photographic films are affected by nuclear radiations, special films which are more sensitive to ionizing radiations rather than light are used to secure better results. The most popular of these films is Kodak "No-Screen" medical x-ray film.

For good resolution--accurate rendition of detail--on the autoradiogram, the specimen and film should be in intimate contact with each other, separated only by the thin Saranwrap. If the specimen and film are not in the closest possible contact, a diffuse image will result. The shape of the specimen also affects resolution; a thick, irregularly-shaped specimen will also cause poor resolution because of the self-absorption and scattering of radiation within the sample.

"No-screen" X-ray film is double-coated; the sensitive emulsion is coated on both sides of the film base, so it does not matter which side faces the source of radiation.

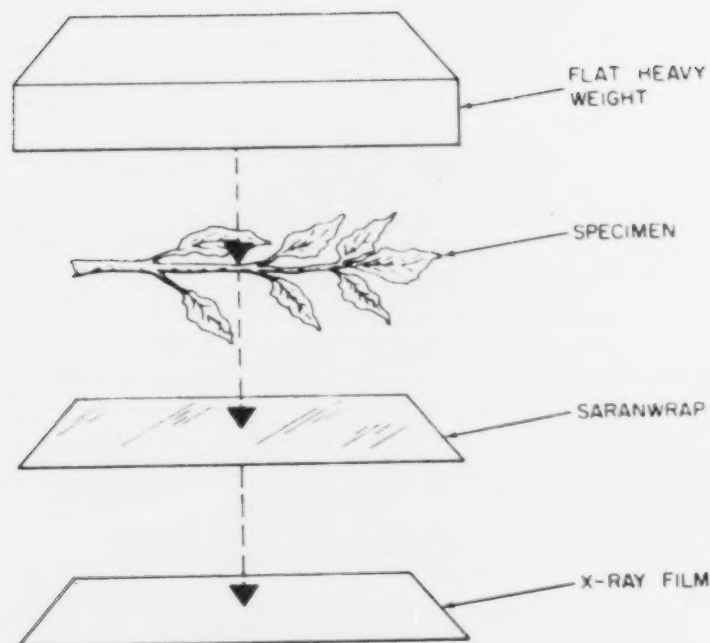


Figure 3. Preparing Autoradiogram of Plant

After positioning the specimen, thin "Saranwrap" and X-ray film, these materials should be kept in close contact for the duration of the exposure by placing a heavy weight upon them or by some other appropriate method. It is important that none of the materials be moved during the exposure.

Calculating the exposure

The exposure time is really a matter of trial and error, but a useful approximation of the exposure time may be calculated with the aid of a Geiger counter. Because of the great variation in specimens, type of isotope, beta energy, to name a few, any calculation of exposure time is only an approximation.

About 10^7 beta particles/cm² are needed to expose X-ray film. Most Geiger counters will count about 3-5% of the beta particles emitted by a source close to the tube. If the count rate of a one cm² specimen is known, a simple relationship may be derived to compute the exposure time.

The count rate of the specimen is obtained by placing a thick lead foil with a one cm² opening over the specimen and placing the probe of the Geiger counter as close as possible to the opening.

A simple formula may be derived to approximate the exposure time:

$$x = \frac{5 \times 10^5}{A} \quad \text{where} \quad x = \text{approximate exposure time in minutes}$$

A = activity in counts/min.
of one cm² of specimen.

EXAMPLE: A one cm^2 area of a leaf which contains P^{32} solution is found to yield an activity of 3000 counts per minute on a Geiger counter. Calculate the approximate exposure time.

$$x = \frac{5 \times 10^5}{A} = \frac{5 \times 10^5}{3000}$$

$$x = \frac{5 \times 10^5}{3 \times 10^3} = 5/3 \times 10^2$$

$$x = 1.67 \times 10^2$$

$$x = 167 \text{ minutes} = 2 \text{ hrs. } 47 \text{ minutes}$$

In this case, exposure times of 2 hrs, 3 hrs, and 3 1/2 hours should be tried with similar samples.

Processing the film

Although "No-screen" x-ray film is more sensitive to ionizing radiations than to light, it will be exposed unless handled in complete darkness or under a Wratten Series 6B safelight filter. The room should be as nearly light-tight as possible with the exception of the safelight. The safelight should be at least 6 feet away from the film processing area.

A table should be arranged near a sink so that three 5 x 7" photographic trays (or a similar size tray or pan) are positioned in the order shown below:

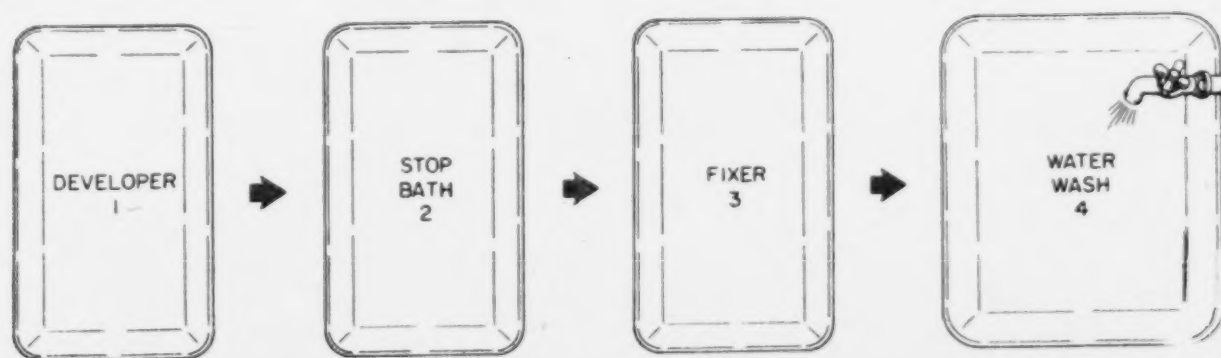


Figure 4. Arrangement of Trays for Processing

Instructions for mixing the developer and fixer solutions are given on the containers. Pour the X-ray developing solution into the first tray, pour a weak solution of acetic acid into the second tray---about 20 ml of white vinegar or one ml of glacial acetic acid per 500 ml of water will do. Pour the x-ray fixer solution into the third tray. The sink should be filled with water as a rinse.

The second and third trays contain a weak acid to arrest the action of the alkaline developer, so at no time should any liquid from these two trays be spilled or added in any way to the first tray. Use the left hand for handling films in the developer, and the right hand for handling films in all the other solutions. This is most important--the acid solutions will neutralize the developer and poor results, if any, will be obtained.

For best results, the temperature of the liquids in the three trays should be maintained as nearly the same as possible and at a value of 68°F (20°C). If this is not done, higher temperatures which may occur in hot weather will cause the gelatine in the emulsion to swell and shrink and displace the silver grains. Higher temperatures can be partially compensated for by shorter development times but tend to produce high contrast along with faster exhaustion of the developer solution. Lower temperatures require longer development times and tend to produce lower contrast.

At 20°C the film should be completely immersed in the developer for 5 to 10 minutes, then quickly transferred to the stop bath for 30-60 seconds, then transferred to the fixer for minutes. The films should be washed in water for 5 to 10 minutes and then hung up to dry in a clean, dust free atmosphere. If a few ml of a mild liquid detergent such as Joy, Swan, Kee, etc., are added to the wash water, the films will dry faster and with fewer water spots.

The processing instructions given above yield a "negative" autoradiogram. - If a positive print is desired, expose the dry negative to photographic paper as shown below:

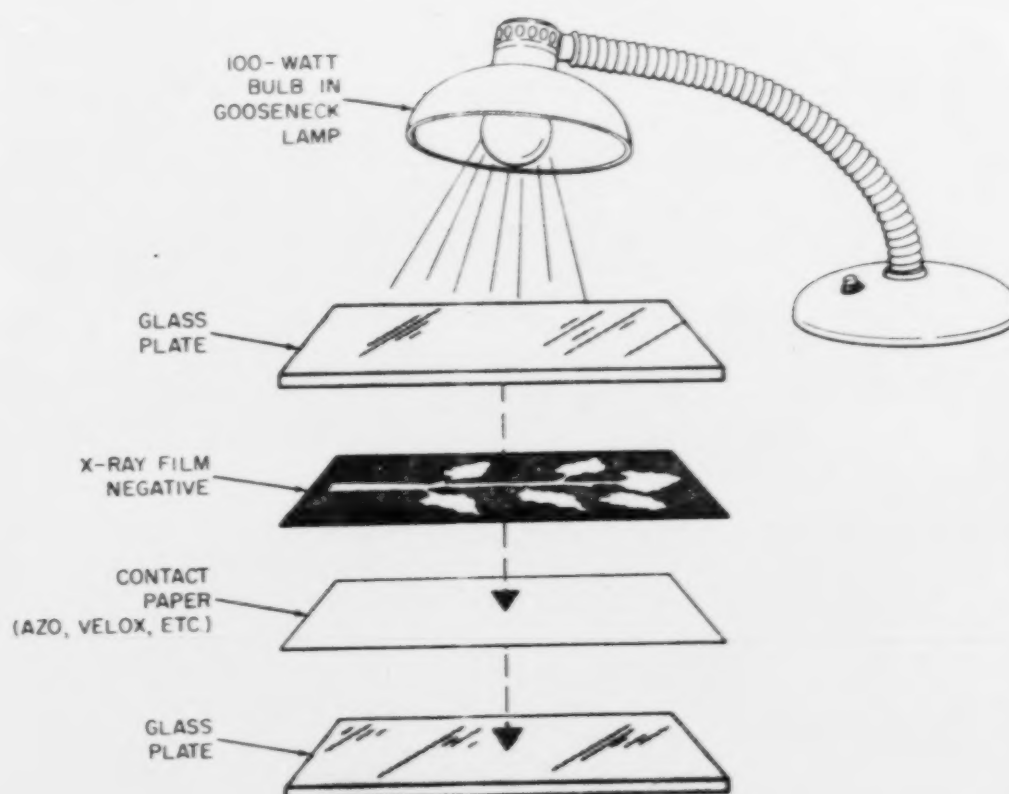


Figure 5. Making a Positive Contact Print from X-ray Film Negative

The Wratten Series 6B safelight must not be used with photographic papers; instead, a Wratten Series 1A, 2, or OA must be used.

Exposure times will vary with the type of photographic paper chosen, and should be determined by trial and error with small bits of the paper.

After the paper has been exposed, it is processed in the same manner as the film, but with a different developer and fixer. Kodak Dektol Developer and Acid Fixer are used. The prints may be dried between sheets of absorbent paper.

SECTION II EXPERIMENTS

Experiment 1

ABSORPTION OF PHOSPHORUS BY PLANTS

PROBLEM

Where does a plant concentrate phosphorus after absorption from a nutrient solution?

INTRODUCTION

Dyes, identifying paint, and tags have been used to follow the paths of animals, birds, the flow of liquids, and in similar applications for centuries. Even early guidons and battle flags were "tracers" utilized to follow the movements of military units.

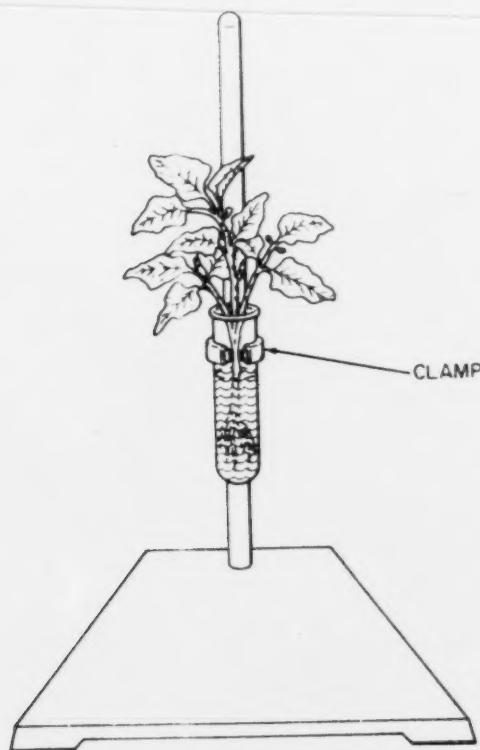


Figure 6. Determining Phosphorus Absorption by Plant

RUBBER GLOVES SHOULD BE WORN IN THIS EXPERIMENT

Radioactive elements undergo the same reactions as non-radioactive elements, and when mixed with the latter, can act as a tracer.

In this experiment, the plant will be immersed into a solution of phosphorus-32. The plant will absorb water and the phosphorus. From an autoradiogram, the concentration and location of the phosphorus will be determined from the darkening of a photographic film.

APPARATUS NEEDED

Developing trays; ring stand; clamp; small test tube; safelight.

MATERIALS NEEDED

Small, young tomato or bean plant; 2 μ c P^{32} solution; developer, stop bath, and fixer.

PROCEDURE

1. Clamp the test tube to the ringstand and add the P^{32} solution. Remember that the stem and roots will be placed in the test tube, so make certain there will be enough room for them without causing liquid in the test tube to overflow.
2. Carefully wash the dirt from the tomato plant without damaging any of the tender roots.
3. Place the tomato plant in the test tube. The roots should be below the surface of the solution. If not, add a minimum amount of water.
4. Allow the plant to remain undisturbed for 24 hours.

If a Geiger counter is available, perform the following steps:

- 5a. Wrap a piece of 1/8" thick lead foil around the test tube.
- 5b. Determine the activity of one of the smaller leaves which has an area close to one cm^2 . If the leaves are too large, cover one with a lead foil which has a one cm^2 opening cut in it.
- 5c. Compute the necessary exposure time from this activity.
6. Cut the plant from the roots at a point just above the liquid level, and prepare an autoradiogram.

If a Geiger counter is not available experiment with different exposure times until a satisfactory autoradiogram is obtained.

CONCLUSIONS

Where does the phosphorus concentrate in the plant? Can you suggest a reason why?

If other isotopes are available, repeat the experiment to determine if they will concentrate in the same areas of the plant as the P^{32} . Especially interesting are Na^{22} and I^{131} .

MAKING INEXPENSIVE FILMSTRIPS

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For many years, the 35 mm filmstrip has been the mainstay of the various multi-media devices available to the science teacher. Recent years have seen the arrival of the overhead projector, the TOPS-type projector, and filmloops to compete for teacher popularity. Yet the 35 mm filmstrip still retains its popularity in terms of actual classroom use. All schools possess some type of filmstrip projector or viewer, and the low cost and versatility of the filmstrip make it especially suitable for classroom use in all disciplines.

Although silent, captioned-types of filmstrips are more popular, the sound type has its adherents. The sound type is not as versatile, though, for the teacher cannot interrupt a tape or a record easily in order to insert comments of his own choosing. Many instructors do not agree in toto with the authors or producers of a filmstrip in either the way in which the filmstrip subject matter is presented or the order of presentation. And rightly so, for it is this type of academic freedom which has led to progress and innovation in teaching.

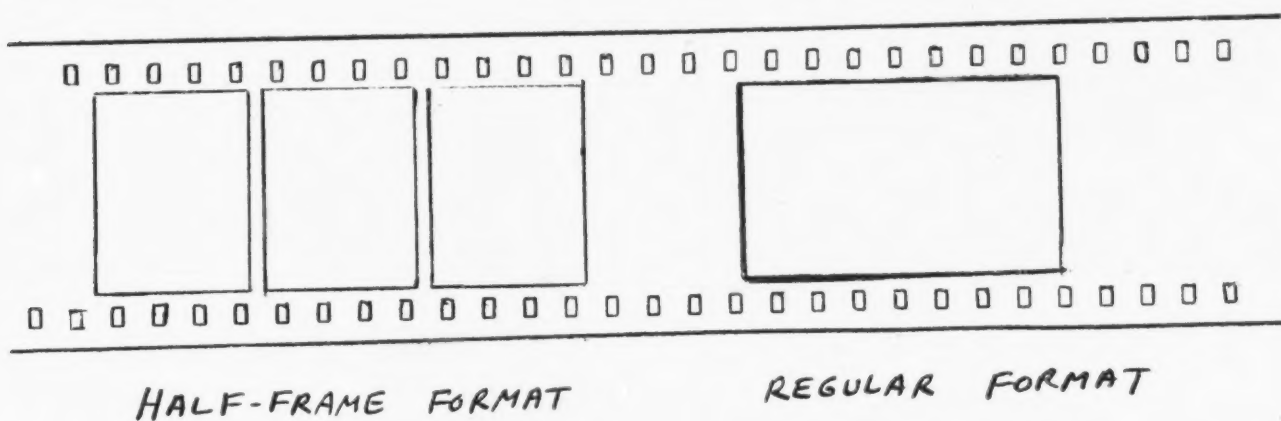
What recourse does the teacher have if he does not agree with a mode of presentation in a particular filmstrip? Four alternatives come to mind:

1. Do not use it;
2. Continue to use it,
3. Cut it apart, remove or rearrange the offending frames, mount the others in half-frame mounts and use in a slide projector;
4. Make your own.

The first choice is somewhat drastic, while the second may not be desirable; the third is quite time-consuming but effective, and the fourth choice is the main thought of this article.

When a commercial filmstrip is produced, the artwork, full-size slides, and typography are photographed in sequence and reduced to what is called "half-frame" dimensions by special cameras on negative color or black-and-white film. This negative color filmstrip is called the master negative; when printed on another negative film material, a positive release

print results. This strip of projectable color film is what we ordinarily call a filmstrip. Essentially, this is the process used for large quantities of release prints.



If only one strip were needed, the artwork, etc., could be shot with positive color film - reversal color film - such as Kodachrome II or Ektachrome-X and developed normally.

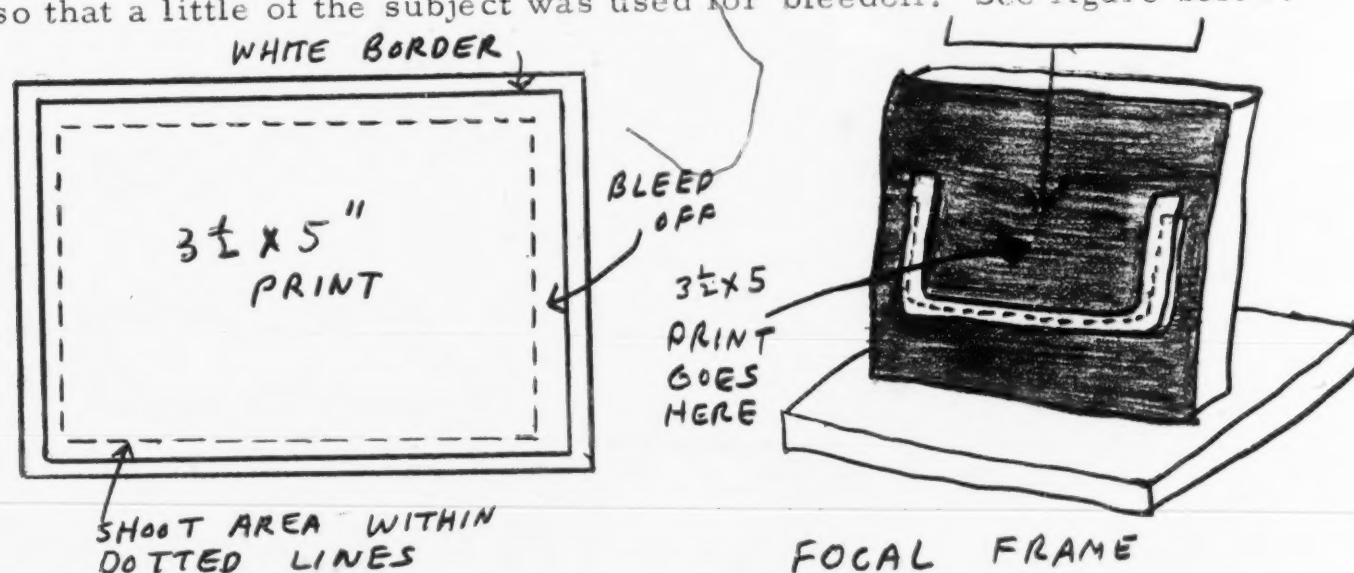
The special cameras used to make a master negative are extremely expensive, and only the biggest film processors and filmstrip manufacturers have them. But half-frame cameras are available to the general public at prices comparable to the other 35 mm cameras. These cameras will give the half-frame or filmstrip format. The only half-frame camera suitable today for home-made filmstrips is the Olympus Pen-F or the Pen FT. These cameras feature through-the-lens viewing by means of the single-lens reflex system. What you see in the finder actually appears on the film, thus eliminating the old problem of parallax present in cameras not of the single-lens reflex type. The Pen FT contains an integral behind-the-lens metering system which aids in computing the exposures.

The subject matter is photographed through the Pen FT in sequence, the film is sent to Kodak (or other processor) with the stipulation "Do not cut, do not mount," and it will be returned as a filmstrip roll--uncut, in the half-frame format ready for projection in a standard 35 mm filmstrip projector.

The above simple explanation does not, of course, include the little touches which provide a professional-looking final product, such as titles and captions. Titles--such as the title of the strip itself, credits, boilerplates, etc.,--may be simply made. Fancy-looking titles may be constructed from movable letter kits such as those sold for home-movie makers. Hand lettered titles made by printing with a felt pen on light yellow construction paper also work very well to provide a stylized, professional look. Light yellow is preferred over other colors - especially white - because it looks much better in the final product. Other colors may look fine to the eye, but end up as an entirely different color when photographed.

The SMIP staff has constructed a small wooden frame which holds 3 1/2" x 5" title cards and color photographs. The title cards are placed in this frame, illuminated at 45° angles from the side by two blue photofloods (when shooting with daylight color film), and photographed in sequence with the Pen FT mounted on a tripod. +2 and +3 auxiliary Portra lenses are mounted on the camera lens in order to allow very close focusing.

The focal frame described above was painted with flat black paint so that any errors in focus or framing would be eliminated. When photographing title cards on the frame, the camera was moved in close enough so that a little of the subject was used for bleedoff. See figure below.



With this frame, artwork, photographs (both black and white and color prints) and titles could be photographed rapidly and easily. An explanatory paragraph, typed on a 3 1/2" x 5" light yellow card with IBM typewriter, photographs very well.

The simplest filmstrip of all to make is the non-captioned type. A little card, a boilerplate card, and "The End" are the only special effects needed; the other scenes are merely shot in sequence. (No mistakes!). When the filmstrip is shown, for example, the teacher can narrate each frame either spontaneously or from a prepared script.

A captioned filmstrip may be constructed in a similar way. The scenes to be used are shot first on Kodacolor film and 3 1/2" x 5" Kodacolor prints are made by the photofinisher. Captions are typed with an electric typewriter on light yellow construction paper, cut apart, and pasted lightly on the Kodacolor prints on the bottom, sides, and top, or wherever desired. The Kodacolor prints are then photographed in sequence on the focal frame with the Pen FT.

Two words of caution are in order here. First, the light yellow paper on a print may sometimes mislead the exposure meter on the Pen FT because it is usually brighter than the rest of the print. With yellow captions, it is wise to overexpose by 1/2 stop to compensate for this effect. Secondly, when photographing the prints in sequence to make the final strip, make certain not to make a mistake on any picture, or the whole strip must be reshot from the beginning.

In shooting the final strip, when all titles, artwork, captioned photos, etc., are assembled, the photographer is faced with a bewildering choice of reversal films from which to choose. The following suggestions are the personal opinions of the author, and do not imply recommendations of any commercial products or films. People have certain preferences for color castes and renditions: where one person will like the color effects given by one film, another person will prefer another version.

<u>Film</u>	<u>ASA Speed</u>	<u>Remarks</u>
Kodachrome II	32	Fine grain. Good fidelity; less contrast.
Kodachrome -X	64	Contrastier than KII, more grainy; usually more pleasing result than KII.
Ektachrome -X	64	May be developed at home with new E-4 process; heightens colors, more vivid; good for blues and greens.
High Speed Ektachrome	160	Similar to EX, but less contrasty, little more grain.
Ansochrome 50, 100, 200, 500		Not recommended; colors not true to author's eye.

In writing a filmstrip script, the primary thought to keep in mind is that a story must be told visually and with continuity. Each frame must logically follow the others, so that any breaks in thought, or loss of continuity, can be avoided. In a sense, a filmstrip is a photographic essay. Even a theory-type filmstrip must be given the same care in regards to continuity as a descriptive or a technique filmstrip. This is why professional filmstrip producers rely heavily on storyboards.

The quality of home-made filmstrips cannot truly approach that turned out by a professional organization utilizing the master negative process. With a little practice, mistakes become fewer and fewer in production, and the final product becomes more professional in nature.

For unusual subjects, unusual or personalized presentations, the home-made filmstrip can fulfill a need. In many cases, the final product is worth the effort.